

**EXPERIMENTAL EDITION**

# **SCHOOL CHEMISTRY**

**CLASS 8**

**TEXT**

FIRST PUBLISHED  
1970

© NATIONAL COUNCIL OF EDUCATIONAL  
RESEARCH AND TRAINING

*All rights whatsoever on this book are strictly reserved and no part of it may be  
reproduced by any process for any purpose without the written  
permission of the copyright owners*

Printed in India  
at Shivaji Press, Secunderabad ( A P. )

# SCHOOL CHEMISTRY

CLASS 8

*TEXT*

*Prepared by*

**N C E R T Chemistry Study Groups**



*Under a grant from*

**THE NATIONAL COUNCIL OF EDUCATIONAL  
RESEARCH AND TRAINING**

**NEW DELHI**



## PREFACE

This book is part of the continuing effort being made by the Study Groups in chemistry set up under the National Council of Educational Research and Training, Ministry of Education, Government of India, in bringing out books in chemistry, in pursuance of the Education Commission's recommendation that chemistry and other sciences be taught as separate disciplines in Indian schools. The material presented in this book is expected to serve as a connecting link between classes 6 and 7, the books for which were brought out earlier and classes 9 and 10 for which the material is under preparation. Attempts have been made to develop the electronic and ionic concepts in a simple style to explain the general principles of chemistry and to interpret various phenomena in the light of these concepts. The periodic table has been included to develop the idea of periodicity of properties of elements and their correlation with the electronic configuration of atoms. Care has been taken to see that necessary descriptive chemistry is included. This modest effort, it is hoped, will help to stimulate the student's interest in the subject.

The members of the Study Groups are grateful to Dr. D.S. Kothari, Chairman, University Grants Commission and Chairman, NCSE, for his encouragement and to the NCERT for providing funds to make the publication of this book possible. They are also grateful to Dr. L.S. Chandrakant, Joint Secretary, Ministry of Education, Government of India and Secretary, NCSE, and to Dr. S.K. Mitra, Joint Director, NCERT, for their encouragement and continued interest in the working of the Study Groups. They are thankful to Prof. J.A. Campbell, Regional Coordinator for science teaching improvement, UNESCO, Prof. Paul R.O' Connor, Staff Associate, NSF science liaison staff, USAID, New Delhi, and Mr. Alan Jackson, Consultant to the Summer Institute in chemistry (1969), Poona, for their perusal of the manuscript and valuable suggestions.

They appreciate the services of Shri G. Sivaram, artist, for his illustrations.

November 20, 1969  
Chemistry Department,  
Osmania University,  
HYDERABAD-7 A.P.

N. V. Subba Rao  
Convener,  
NCERT Study Groups in chemistry.



# DIRECTORS

AND

# MEMBERS OF THE STUDY GROUPS

## CONVENER

Prof. N V. SUBBA RAO;  
Osmania University,  
HYDERABAD.

## DIRECTORS

Prof H J Arnika,  
University of Poona,  
POONA.

Prof. R.D. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Prof. R.C Paul,  
Panjab University,  
CHANDIGARH.

Rev. Father L.M Yeddanapalli, S.J.,  
Loyola College,  
MADRAS

## MEMBERS

Dr, Ashok Kumar Sharma,  
Panjab University,  
CHANDIGARH.

Dr. S. Gireesan,  
Loyola College,  
MADRAS.

Dr A. V Khedekar,  
University of Poona,  
POONA.

Dr. R P. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Sri M.D. Siddhanti,  
Osmania University,  
HYDERABAD.

Sri B. Subrahmanyam,  
Osmania University,  
HYDERABAD.

Sri Venkobacharya Upadhyaya,  
Osmania University,  
HYDERABAD.

## MEMBERS ASSOCIATED WITH THE STUDY GROUPS

Sri S.V. Appa Rao,  
New Science College,  
HYDERABAD.

Sri S A. Basha,  
New College,  
MADRAS.

Sri G.Y. Dixit,  
N M.V High School,  
POONA.

Sri L.H. Gadgil,  
S P. College,  
POONA.

Dr. Gurdev Singh,  
Panjab University,  
CHANDIGARH.

Kum. K. Hema,  
Sacred Hearts High School,  
MADRAS.

Sri D.V. Kelkar,  
M.E.S College for Girls, (D.G.),  
POONA.

Sri P.S. Krishnaswamy,  
R K.M. High School,  
MADRAS.

Sri S.R. Kulkarni,  
New English School,  
POONA,

Dr. M.L. Lakhanpal,  
Panjab University,  
CHANDIGARH.

Dr. K.C. Malhotra,  
Panjab University,  
CHANDIGARH

Prof. B.N. Mankad,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Prof. K S. Narang,  
Panjab University,  
CHANDIGARH.

Dr. C.K. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR

Dr. K.C. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Dr. M.R. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Prof. S. R. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Sri V.M. Patel,  
Sardar Patel University,  
VALLABH VIDYANAGAR.

Dr. V K. Phansalkar,  
University of Poona,  
POONA

Prof B.R. Puri,  
Panjab University,  
CHANDIGARH

Dr. N. Venkatasubramanian,  
Vivekananda College,  
MADRAS.



## CONTENTS

CHAPTER		PAGE
1.	ACTION OF HEAT ON SUBSTANCES ...	1
2.	ELECTRICAL NATURE OF MATTER ....	7
3.	STRUCTURE OF ATOM ...	15
4.	HOW ATOMS COMBINE . ...	26
5.	CHEMICAL EQUATIONS ..	33
6.	HYDROGEN, OXYGEN AND WATER ...	43
7.	SOME CHEMICAL FAMILIES ....	61
8.	PERIODIC CLASSIFICATION OF ELEMENTS ...	71

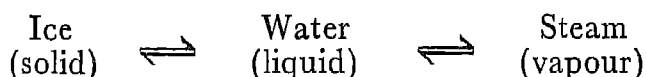


## ACTION OF HEAT ON SUBSTANCES

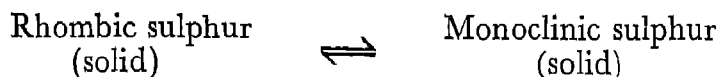
Heat produces many changes in substances such as change of state, combination, decomposition and rearrangement of atoms in molecules. In this chapter we shall learn in some detail the effect of heat on a few substances.

### 1.1 Heat producing no change in composition

We know that ice changes to water on heating. On further heating water changes to steam.



Such changes of state from solid to liquid, liquid to vapour are called *phase changes*. Each state of matter, which is homogeneous by itself, and is separated from another by a boundary is known as a *phase*. We have yet another example of phase change—a solid is converted into another solid: rhombic sulphur is converted into monoclinic sulphur when the former is heated to a temperature of about 96°C.



These changes can be reversed\*. There is no change in the chemical composition of the substance. Water, whether it exists as solid, liquid or vapour has the same composition represented by the formula  $\text{H}_2\text{O}$ .

Let us see what happens when a piece of ice, say, at  $-10^\circ\text{C}$ , is heated continuously. The temperature of the ice piece slowly rises until it reaches  $0^\circ\text{C}$ . The temperature remains constant and does not rise further even though the supply of heat is continued. But we notice that the ice

---

\* In many cases the changes may take place slowly.

starts melting and that ice and water exist together. When all the ice melts the temperature commences to rise again and continues to do so until it reaches  $100^{\circ}\text{C}$ . At this temperature water starts boiling and is converted into steam. Water and steam exist together. As long as conversion of water to steam takes place, the temperature remains constant. When all the water is vapourised, the temperature rises once again as the supply of heat is continued. The change of temperature with time, as heat is continuously supplied, is illustrated in figure 1.1.

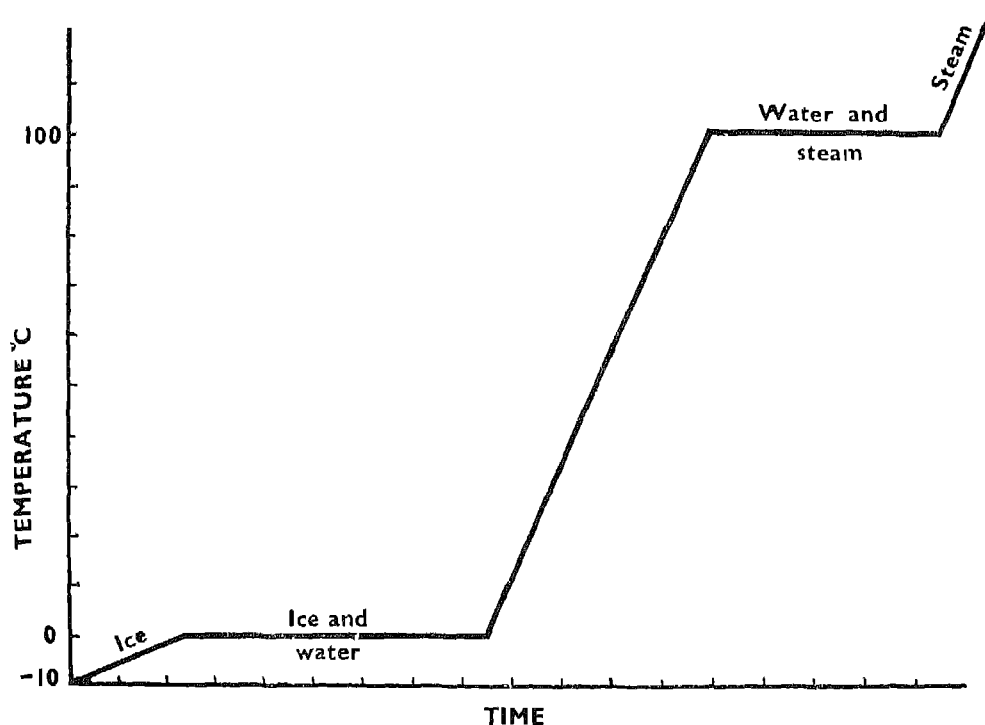


Fig. 1.1 Phase changes (schematic).

When steam is cooled, the changes take place in the reverse direction. Similar phase changes occur in other substances also; but the temperature at which they take place is characteristic of the substance.

Why does temperature remain constant when phase changes take place?

We studied that matter is made up of tiny particles which possess kinetic energy. They are held together by forces of attraction. The heat supplied is used up in two ways: (i) to increase the velocity of

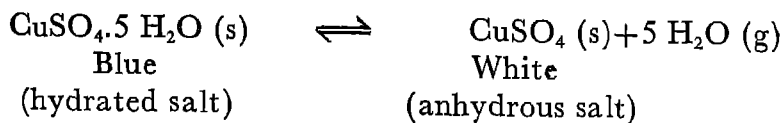
the particles and thereby their kinetic energy, (ii) to overcome the forces of attraction between the particles. When we start heating a substance the particles pick up energy and move faster and the increase in thermal agitation is reflected as rise in temperature. Once the particles become sufficiently energetic, they overcome the forces of attraction and fly apart. At this stage heat is used in overcoming the forces of attraction and not in increasing the thermal agitation. That is why the temperature remains constant during the phase changes.

The action of heat on camphor is different from that on ice. Under ordinary conditions, camphor, a solid, vapourises directly on heating, instead of being converted to a liquid. Cooling the vapour gives back solid camphor. The conversion of a solid directly into vapour without passing through the liquid phase is called sublimation. Iodine, naphthalene and ammonium chloride are other examples of substances which sublime.

## 1.2 Heat producing change in composition

In the changes discussed above, we find that there is no change in either the mass or the composition of the substance during heating. Let us now see what happens when substances like copper (II) sulphate, barium chloride and copper (II) carbonate are heated.

Crystalline copper (II) sulphate,  $\text{CuSO}_4 \cdot 5 \text{H}_2\text{O}$ , is blue. A mole of this compound contains five moles of water of crystallization. A salt like this which contains water of crystallization is called a hydrated salt. On heating, it loses water and becomes a white powder,  $\text{CuSO}_4$ .



The white powder becomes blue on the addition of water showing that the change is reversible. The composition of the blue crystals is different from the composition of white powder. Cobalt (II) chloride behaves in a similar manner. The hydrated salt,  $\text{CoCl}_2 \cdot 2 \text{H}_2\text{O}$ , is pink. On heating, it loses water and turns blue. The pink colour is restored on adding water to the anhydrous salt.

When hydrated barium chloride,  $\text{BaCl}_2 \cdot 2 \text{H}_2\text{O}$ , is heated it loses water of crystallization and forms the anhydrous salt,  $\text{BaCl}_2$ .

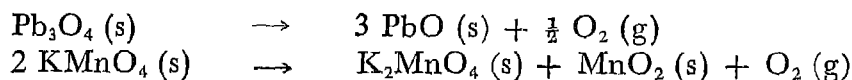


Due to the escape of water, the salt loses weight. Other hydrated salts also behave similarly.

Greenish blue copper (II) carbonate, on heating, gives black copper (II) oxide and carbon dioxide.



The carbon dioxide escapes to the atmosphere. Similarly, when red lead,  $\text{Pb}_3\text{O}_4$  and potassium permanganate,  $\text{KMnO}_4$ , are heated, oxygen escapes to the atmosphere and there will be a loss in the weights of the substances taken.



Thus, substances which yield gaseous products on heating lose weight.

There are, however, substances which gain weight when heated in air. For example, magnesium, when heated burns with a bright light and forms a white powder which is largely magnesium oxide.\* The weight of the magnesium oxide is greater than that of magnesium. The increase in weight is due to the combination of magnesium with atmospheric oxygen.

*Example 1* What will be the percentage loss in weight when magnesium carbonate, is strongly heated? The equation for the decomposition is



Formula weight of magnesium carbonate = 84.0

84.0 g of magnesium carbonate, when heated, yields 44.0 g of carbon dioxide

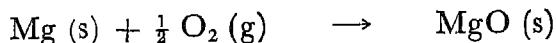
$$\therefore 100.0 \text{ g yield } \frac{44.0}{84.0} \times 100.0 = 52.4$$

Percentage loss in weight = 52.4

---

\* Part of the magnesium combines with nitrogen to form magnesium nitride,  $\text{Mg}_3\text{N}_2$ .

*Example 2* Calculate the gain in weight when 1.0 g of magnesium is completely burnt in air to give magnesium oxide.



24.0 g of magnesium, when completely burnt in air, gives 40.0 g of magnesium oxide

$$\therefore 1.0 \text{ g gives } \frac{40.0}{24.0} \text{ g} = 1.7 \text{ g}$$

$$\text{Gain in weight} = 1.7 - 1.0 = 0.7 \text{ g}$$

There are a number of substances which, on heating, change in appearance, but do not undergo change in weight, e.g. zinc oxide, ZnO. It is white when cold but becomes yellow when heated. The composition remains the same.

When urea is heated, ammonium cyanate is formed.



The two substances are quite different and have different properties, though the number and kind of atoms in each are the same. The difference in properties is due to the different arrangement of the atoms in the substances.

Thus, heat produces changes in phase, changes in composition and changes in the arrangement of atoms in molecules.

## QUESTIONS

1. What is a phase change? Give examples.
2. Describe an experiment to prove that when magnesium burns, it combines with the oxygen of the atmosphere.
3. Does iron gain or lose weight on rusting? Explain.
4. When one ton of limestone is heated, does the resulting quicklime weigh more or less? Why?
5. Fill in the blanks :
  - (a) When.....copper sulphate crystals are heated in a test tube, they turn .....and at the same time, .....are seen on the cooler parts of the test tube.

(b) When red is heated the gaslead evolved is. ....

(c) Conversion of water to steam is a. ....change.

(d) Hydrated baium chloride, when strongly heated, gives the.. ..salt.

6. Match the appropriate phrases from Group B with the words in Group A.

A

1. Phase change
2. Decomposition
3. Composition
4. Rearrangement

B

- (a) Burning of magnesium
- (b) Heating urea
- (c) Heating zinc oxide
- (d) Heating lead nitrate
- (e) Sublimation of camphor



## ELECTRICAL NATURE OF MATTER

We have already studied that matter contains electrically charged particles. In this chapter the electrical nature of matter is further investigated by studying the effect of heat and light on metals and also of electricity on gases, solids and liquids

### 2.1 Thermionic emission

Thomas Edison found, in 1883, that a stream of negatively charged particles flows through an evacuated discharge tube fitted with a heated cathode and a cold anode. A form of the apparatus used for this purpose is shown in figure 2.1.

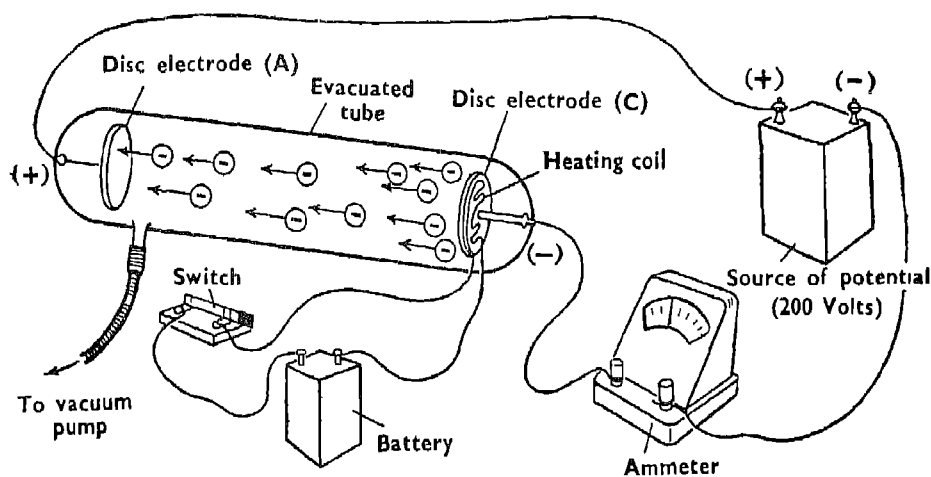


Fig. 2.1 Thermionic emission

The apparatus consists of a glass tube fitted with two disc electrodes A and C. The electrode C on the right is the cathode, to which a heating coil is attached. The cathode is heated by passing a current from a battery through the coil. The external circuit between the anode A and cathode C includes an ammeter and a source of potential. The glass tube is evacuated and the potential applied. The ammeter does not show any deflection as long as the cathode is cold. But on heating

the cathode the ammeter shows a deflection, indicating flow of charge. If the hot electrode is made positive and the cold one negative by reversing the connections, no charge flows through the gas. Evidently charge flows only from the hot cathode to the cold anode.

Since a hot cathode is essential for the charge to flow, it may be concluded that high temperature loosens and releases some negatively charged particles from the cathode. This phenomenon is known as *thermionic emission* and is made use of in the electronic valves used in radio sets.

## 2.2 Photoelectric emission

Metals like lithium, sodium and potassium are found to emit negatively charged particles from their surfaces when light of suitable energy falls on them. This phenomenon is known as *photoelectric emission* and was discovered by Hertz in 1887.

## 2.3 Gas discharge

Many experiments were carried out to study the conduction of electricity by gases. A glass tube about 50 cm long and containing air (or any other gas) is fitted with electrodes as shown in figure 2.2. When a potential of

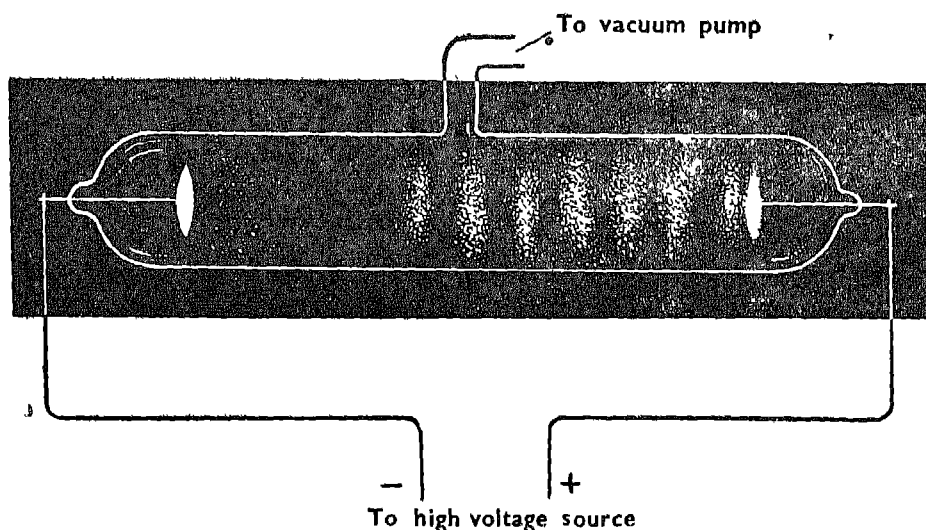


Fig. 2.2 Discharge of electricity through a gas.

about 20,000 volts is applied between the electrodes no electricity is conducted between the electrodes.

If, however, some of the gas is pumped out of the tube (i. e. if the pressure of the gas is reduced) electricity begins to flow and the gas glows (emits light). As the pressure is reduced further a dark space appears near the cathode and alternate bright and dark regions are observed in the rest of the tube. At still lower pressures the dark space increases in size until it fills the whole tube. No light is now emitted by the gas. Instead the wall of the tube opposite the cathode glows with a faint greenish light.

The glow of the glass is due to its being hit by rays coming from the cathode. Experiments show that these rays (called *cathode rays*) consist of negatively charged particles. The same kind of particles is obtained no matter what the gas in the discharge tube is or what metal is used as the cathode.

It was shown (mainly as a result of J. J. Thomson's work) that the negatively charged particles emitted in all the three phenomena—thermionic emission, photoelectric emission and gas discharge—are identical. These particles have been called *electrons*. It is now known that the electron is an essential constituent of all matter. It possesses a negative electric charge. No particle with a smaller electric charge than that of the electron has been found. Therefore the charge of the electron is taken as the *unit* of electric charge. The electron has a mass which is about  $1/1837$  of the mass of a hydrogen atom. For most purposes it is regarded as having a negligible mass.

## 2.4 Electrical conduction in metallic substances

Metals are good conductors of electricity. On the other hand substances like cardboard, plastic, rubber, etc., are poor conductors. These are known as *insulators*. Metals contain some electrons which have considerable freedom of motion. When an electric potential is applied across a metallic wire the mobile electrons move from atom to atom. This flow of electrons constitutes the passage of electricity through metals. Conduction in metals is known as *metallic* or *electronic* conduction.

In the case of insulators the electrons are greatly restricted in their movement. Hence electricity does not flow easily through insulators.

## 2.5 Electrical conduction in liquids and solutions

Salts like sodium chloride, lead bromide, copper (II) chloride, etc., are non-conductors in the solid state. In the molten state, however, they conduct electricity. Their aqueous solutions also conduct electricity. The passage of a direct electric current through molten salts or their aqueous solutions results in their decomposition. This decomposition is called *electrolysis*. The substances are called *electrolytes*.

Liquids like kerosene, benzene or aqueous solutions of substances like sugar do not conduct electricity. Such substances are called *non electrolytes*.

In an ionic crystal like sodium chloride the sodium ions,  $\text{Na}^+$ , and the chloride ions,  $\text{Cl}^-$ , are firmly held in place by the neighbouring ions. When a crystalline salt is melted the ions move about with considerable freedom. In the aqueous solution of the salt also, the ions are free to move about. When two electrodes are placed in the molten salt or its aqueous solution, and a potential applied across them, electricity is carried by the ions. Thus the electrical conduction in a molten electrolyte or its aqueous solution is different from metallic conduction. It is due to the motion of the ions in the molten liquid or solution. Such conduction is called *ionic* or *electrolytic* conduction.

During the electrolysis of molten sodium chloride, the sodium ions,  $\text{Na}^+$ , move to the cathode. Electrons are made available at the cathode by the source of electricity. Each sodium ion,  $\text{Na}^+$ , takes up an electron from the cathode to form a sodium atom.



The chloride ion,  $\text{Cl}^-$ , moves to the anode, gives up an electron and forms a chlorine atom.



The electron released flows back to the source of electricity (Fig. 2.3). Notice that the source of electricity does not create electrons. It merely moves them from one place to another. It acts as an electron pump.

Two chlorine atoms combine to form a molecule of chlorine.

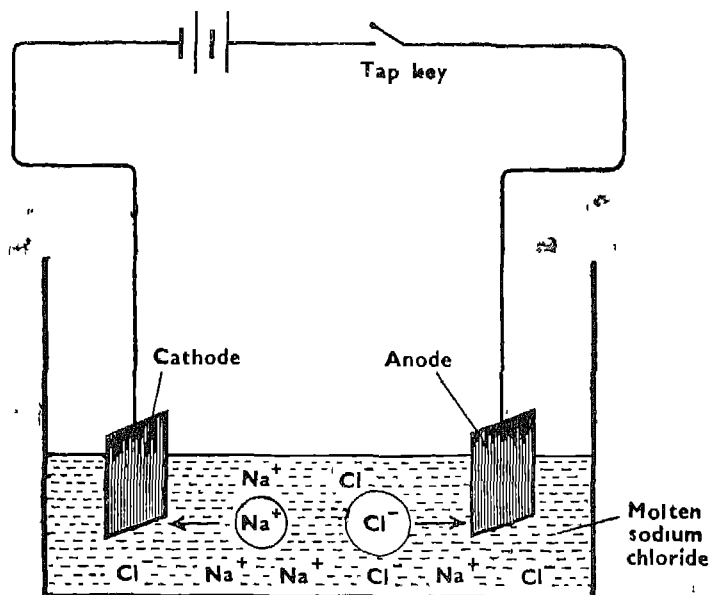
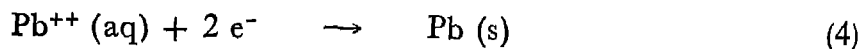


Fig. 2.3 Electrolytic conduction

The reactions (1) and (2) are called electrode reactions. The former is the *cathode reaction* and the latter the *anode reaction*.

In a similar manner, electrolysis of molten lead bromide can be represented by the following electrode reactions :

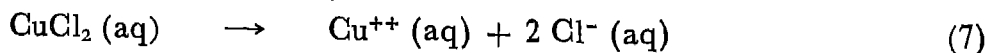
At the cathode



At the anode

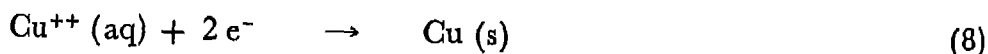


An aqueous solution of copper (II) chloride contains copper (II) ions,  $\text{Cu}^{++}$ , and chloride ions,  $\text{Cl}^-$ .



The reactions at the electrodes are :

At the cathode



At the anode



## 2.6 Relation between quantity of electricity and the amount of material deposited

In the electrolysis of copper (II) chloride the amount of copper deposited at the cathode is directly proportional to the quantity of electricity (i. e. the product of the current strength and the time of passage of the current).

In the case of a univalent ion (i.e. an ion like  $\text{Na}^+$ ,  $\text{Ag}^+$ ,  $\text{Cl}^-$ , bearing a single charge) the quantity of electricity required to deposit one mole ( $6.0 \times 10^{23}$  atoms) of the substance is found to be equal to 96,500 coulombs\*. This quantity of electricity is called the *Faraday*.

1 Faraday = 96,500 Coulombs

Similarly the quantity of electricity required to deposit one mole of a divalent ion (i.e. an ion like  $\text{Pb}^{++}$ ,  $\text{SO}_4^{--}$ ,  $\text{Cu}^{++}$ , bearing two charges) is two faradays. Table 2.1 indicates the quantity of electricity required to deposit one mole each of some metals.

TABLE 2.1  
Relation between electricity and mass

One mole of	Weight in grams	Ion	Quantity of electricity in faradays
Silver	108.00	$\text{Ag}^+$	1
Copper	63.54	$\text{Cu}^{++}$	2
Aluminium	27.00	$\text{Al}^{+++}$	3
Sodium	23.00	$\text{Na}^+$	1
Lead	207.00	$\text{Pb}^{++}$	2

\* Coulomb is the quantity of electricity carried when a current of strength one ampere flows for one second.

*Example 1* Calculate the weight of chlorine produced when a current of 10.0 amperes is passed through molten sodium chloride for 5 minutes.

$$\begin{aligned}
 \text{Quantity of electricity} &= \text{amperes} \times \text{seconds.} \\
 &= 10.0 \times 5.0 \times 60.0 \text{ coulombs} \\
 &= 3,000.0 \text{ coulombs} \\
 &= \frac{3,000.0}{96,500} \text{ faraday} \\
 &= 0.0311 \text{ faraday}
 \end{aligned}$$

The anode reaction is



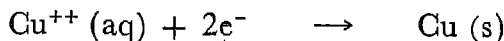
2 faradays of electricity produce 1 mole of chlorine.

$$\begin{aligned}
 \therefore 0.0311 \text{ faraday will produce } &\frac{0.0311 \times 1}{2} \text{ mole of chlorine.} \\
 &= 0.0156 \text{ mole of chlorine}
 \end{aligned}$$

$$\begin{aligned}
 \text{Number of grams of chlorine produced} &= \frac{0.0156 \times 71.92}{1} \\
 &= 1.1 \text{ g}
 \end{aligned}$$

*Example 2* How long does it take to deposit 1.0 g of copper from an aqueous solution of copper (II) sulphate at a constant current of strength 3.0 amperes?

The cathode reaction is



Hence, 2.0 faradays of electricity are required to deposit 1.0 mole (63.54 g) of copper.

$\therefore$  The deposition of 1.0 g of copper requires

$$\begin{aligned}
 &\frac{2.0}{63.54} \text{ faradays} \\
 \text{or } &\frac{2.0 \times 96500.0}{63.54} \\
 &= 3038.0 \text{ coulombs of electricity.}
 \end{aligned}$$

$$\begin{aligned}
 \text{Quantity of electricity} &= \frac{\text{current strength}}{\text{in amperes}} \times \frac{\text{time in}}{\text{seconds}} \\
 3038.0 &= \frac{3.0}{1} \times t \\
 t &= \frac{3038.0}{3} = 1013.0 \text{ seconds} \\
 &= 16 \text{ minutes } 53 \text{ seconds}
 \end{aligned}$$

∴ Time required for the deposition of 1.0 g of copper  
= 16 minutes 53 seconds.

What we have studied here establishes the electrical nature of matter and that electrons are present in all matter. But an atom is electrically neutral. Hence, there must be some positively charged particles in the atom to balance the negative charge of the electrons. What is the nature of these positively charged particles? We shall find answer to this question in the next chapter.

## QUESTIONS

1. Classify the following substances into conductors and insulators :
  - a. Mercury
  - b. Glass
  - c. Graphite
  - d. Aluminium
  - e. Cardboard
2. Classify the following into electrolytes and nonelectrolytes :
  - a. Alcohol
  - b. Water
  - c. Aqueous solution of potassium chloride
  - d. Aqueous solution of copper (II) sulphate
  - e. Aqueous solution of sugar
3. Under what conditions does a gas conduct electricity?
4. What is the smallest unit of negative charge?
5. In the electrolysis of aqueous copper (II) bromide 0.500 g of copper is deposited at one electrode. How many grams of bromine are formed at the other electrode?
6. What is a 'faraday'?
7. Calculate the quantity of electricity (in coulombs) required to deposit
  - (a) 0.12 mole of copper and from copper (II) nitrate solution,
  - (b) 0.12 mole of silver from silver nitrate solution.



## STRUCTURE OF ATOM

We learnt in the previous chapter that electrons are present in all forms of matter and that they are negatively charged. The atom, being electrically neutral, must, therefore, contain particles which balance the negative charge of the electrons. In this chapter we shall learn which other particles are present in the atom and about their nature. We shall also learn how these different particles are arranged in the atom.

### 3.1 The positive ions (Positive rays)

Experiments first performed by Goldstein in 1886 suggested that positive particles are also formed in discharge tubes. A form of the apparatus used by him is shown in figure 3.1.

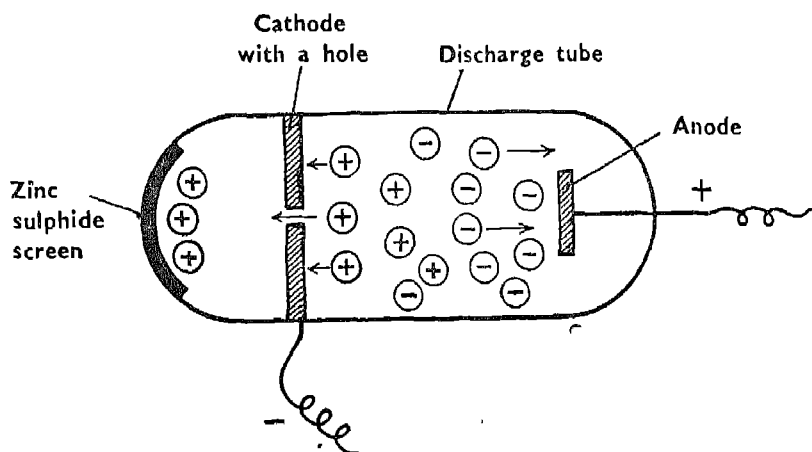


Fig. 3.1 Discharge tube used for studying positive particles

The electrode on the right is positive; the one on the left is negative and consists of a metal with a hole in it. A detecting screen is placed to the left of cathode. When a potential is applied a line of light appears on the screen and is deflected by magnetic and electric fields. The deflection is opposite to that produced in the case of cathode rays. This indicates that a stream of positive rays exists to the left of the cathode. These rays are found to consist of positively charged particles and have the same magnitude of electric

charge (or a simple multiple) as that of the electron. Unlike the electrons the mass of the positive particles depends on the nature of the gas in the discharge tube. When hydrogen is the gas in the discharge tube, the positive particle formed has a mass very close to the *unit of atomic mass* (1 amu). It has the smallest positive charge ever found and is therefore taken as the unit of positive charge. This particle is known as *proton*.

### 3.2 Radioactivity

Henri Becquerel, in 1896, found that the compounds of uranium affected photographic plates even when wrapped in black paper. He concluded that this was due to emission of invisible rays by the uranium salts. He called this phenomenon *radioactivity*. Later work by Marie and Pierre Curie and others showed that besides uranium, other elements like polonium, radium, thorium, etc., were also capable of emitting these rays. Such elements are called *radioactive elements*.

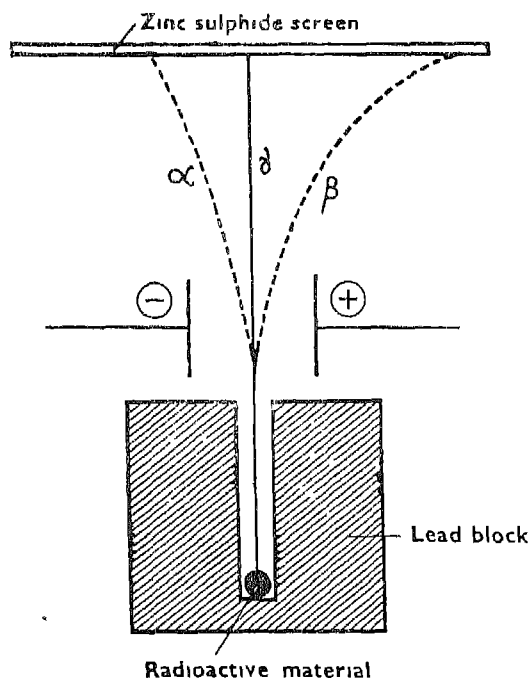


Fig. 3.2 Radioactive emanations.

Experiments conducted by Rutherford and others show that the rays emitted by radioactive elements are of three types. These rays consist of ;

- i) positively charged particles called alpha ( $\alpha$ ) particles. These are helium atoms from which two electrons have been removed, i.e. they are doubly charged helium ions,  $\text{He}^{++}$ .
- ii) negatively charged particles called beta ( $\beta$ ) rays. These are identical with electrons; and
- iii) gamma ( $\gamma$ ) rays, similar to X-rays, but possessing much higher energy. The behaviour of these different rays in a magnetic or electric field is indicated in figure 3.2.

The phenomenon of radioactivity indicates that atoms of some elements are unstable.

### 3.3 The atomic nucleus

Further information about the structure of atom was obtained as a result of Rutherford's experiments. His work clearly revealed the manner in which different particles are arranged in an atom.

Figure 3.3 shows details of Rutherford's experiment. A beam of alpha particles emitted by a radioactive element (polonium), are made to bombard a thin metal foil (gold) placed in the path of the beam. The

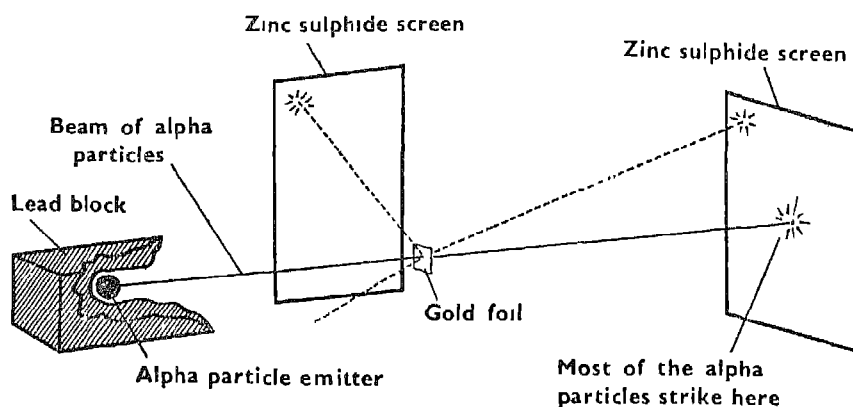


Fig. 3.3 Rutherford's experiment for studying the scattering of alpha particles by metal foil.

alpha particles are detected by a screen coated with zinc sulphide. A flash of light is produced when an  $\alpha$ -particle strikes the screen. Rutherford found that most of the alpha particles went straight through the foil. Some were deflected at large angles. A few were actually deflected back along their path.

From these observations Rutherford concluded that almost the entire mass of the atom and all the positive charges are concentrated in a very small region which he called the *atomic nucleus*. From his observations Rutherford estimated that the diameter of the nucleus is about 10,000 times smaller than the diameter of the atom as a whole.

### 3.4 Atomic number

The work of Moseley showed that atoms of a given element have a whole number of positive charges, i. e. protons, in their nuclei. This number, characteristic of each element, is known as the *atomic number* of the element. Thus, the atomic number of an element is the number of protons in the nucleus. In a neutral atom the atomic number is also equal to the number of electrons outside the nucleus. The atomic numbers of some elements are given in Table 3.1.

TABLE 3.1  
Atomic numbers of some elements

Element	Symbol	Atomic number
Hydrogen . .	H	1
Helium ...	He	2
Lithium .	Li	3
Beryllium ..	Be	4
Boron ...	B	5
Carbon ...	C	6
Nitrogen ...	N	7
Oxygen ...	O	8
Fluorine ...	F	9
Neon ...	Ne	10
Sodium ...	Na	11

### 3.5 The neutron

From Table 3.1 we find that the atomic number of helium is 2. That is, the helium atom has 2 protons in its nucleus. The mass due to

these 2 protons is approximately 2 amu. But the atomic mass of helium is known to be approximately 4 amu. Since the mass of an atom is almost entirely due to the nucleus, it is evident that the protons alone cannot account for atomic mass. There should be some other particles present in the nucleus.

Experiments have since shown that the atomic nucleus contains, besides protons, another type of particles called *neutrons*. As their name suggests, neutrons do not possess any electric charge. The discovery of the neutrons proved to be difficult. Though their existence was postulated in 1920, they were actually discovered in 1932, by Chadwick. As they do not contain any charge they are not deflected by electric or magnetic fields, and so remained undetected for a long time. The neutron has approximately the same mass as that of a proton, i.e. one atomic mass unit.

The discovery of neutrons has greatly simplified the attempts to account for the mass of an atom. The nucleus of every atom (except that of hydrogen which consists of a single proton) contains protons and neutrons, which together are responsible for the mass of the atom. The atomic mass of an element is approximately the sum of the masses of the protons and neutrons present in it. The sum of the number of protons and neutrons is known as the *mass number*.

Table 3.2 gives the atomic masses of some elements in terms of the number of protons and neutrons present in the nuclei of their atoms.

TABLE 3.2  
Atomic masses of some elements

Element	Atomic number (Z)	Particles in the nucleus		Total No. of protons and neutrons	Atomic mass
		Protons	Neutrons		
Hydrogen	1	1	0	1	1.00
Helium	2	2	2	4	4.00
Carbon	6	6	6	12	12.00
Sodium	11	11	12	23	23.00

### 3.6 Isotopes

All atoms of a given element have the same atomic number. However, experiments show that atoms of a single element sometimes have

different mass numbers. For example, hydrogen consists of atoms having three different mass numbers: 1, 2 and 3. As all the three atoms have the same atomic number, i.e. 1, they contain the same number of protons, but different numbers of neutrons. Atoms of an element which have the same atomic number but different mass numbers (and so different atomic masses) are called *isotopes*. The two heavier isotopes of hydrogen, with mass numbers 2 and 3, are known as deuterium and tritium respectively. The deuterium nucleus contains one proton and one neutron, while the tritium nucleus contains one proton and two neutrons.

The isotopes of hydrogen are represented by writing their symbols  ${}_1\text{H}^1$ ,  ${}_1\text{H}^2$  and  ${}_1\text{H}^3$ .

The number on the top right indicates the mass number while the one at the bottom left indicates the atomic number. Table 3.3 gives the composition of the isotopes of a few elements.

TABLE 3.3  
Isotopes of hydrogen, carbon, chlorine and oxygen

Element	Isotope	Mass No.	At. No.	Number of		
				Protons	Neutrons	Electrons (outside the nucleus)
Hydrogen	${}_1\text{H}^1$	1	1	1	0	1
	${}_1\text{H}^2$ (deuterium)	2	1	1	1	1
	${}_1\text{H}^3$ (tritium)	3	1	1	2	1
Carbon	${}_6\text{C}^{12}$	12	6	6	6	6
	${}_6\text{C}^{13}$	13	6	6	7	6
	${}_6\text{C}^{14}$	14	6	6	8	6
Chlorine	${}_{17}\text{Cl}^{35}$	35	17	17	18	17
	${}_{17}\text{Cl}^{37}$	37	17	17	20	17
Oxygen	${}_8\text{O}^{16}$	16	8	8	8	8
	${}_8\text{O}^{17}$	17	8	8	9	8
	${}_8\text{O}^{18}$	18	8	8	10	8

Most of the elements are actually mixtures of isotopes. The atomic masses that chemists measure are, in fact, average masses of the mixture of isotopes of each element. This is why the atomic masses of some elements are not whole numbers.

### 3.7 Arrangement of electrons around the nucleus

From a large body of experimental data Neils Bohr gave a simple model of the arrangement of electrons around the nucleus\*. According to Bohr, the electrons move in shells around the nucleus of the atom. But the electrons are restricted in their movement to a fixed number of possible shells. These shells differ in their average distance from the nucleus. Starting with the one closest to the nucleus these shells are numbered 1, 2, 3 etc. These numbers are called *principal quantum numbers* ( $n$ ). The shells are sometimes designated by letters as follows :

Principal quantum No. ( $n$ )	1	2	3	4
Letter designation of the shell	K	L	M	N

Only a specified number of electrons can occupy any given shell. The maximum number of electrons present in any given shell is given by  $2n^2$  where  $n$  is the principal quantum number of the shell. The maximum *electron population* for different shells is shown in Table 3.4.

TABLE 3.4  
Maximum electron population of different shells

Shell		Principal quantum number ( $n$ )	Maximum number of electrons ( $2n^2$ )
K	...	1	$2 \times (1)^2 = 2$
L	...	2	$2 \times (2)^2 = 8$
M	...	3	$2 \times (3)^2 = 18$
N	...	4	$2 \times (4)^2 = 32$
O	...	5	$2 \times (5)^2 = 50$

Starting with the shell for which the principal quantum number is 1, the electrons are arranged in different shells, according to the capacity of each shell. Let us consider the structure of the atoms of a few elements. The hydrogen atom (atomic number 1) has one electron outside the nucleus. This electron occupies the K shell. The next element, helium, with atomic number 2, has 2 electrons, which also occupy the K shell. Now the K shell is filled to capacity.

\* The model suggested by Bohr has been modified in succeeding years by scientists. The present day model of the atom is vastly different from the one suggested by Bohr. You will learn more about the structure of the atom in higher classes.

In the third element lithium (atomic number 3), two electrons occupy the K shell, and the third goes into the L shell. The L shell can take a maximum of 8 electrons. Therefore, in the seven elements succeeding lithium, the additional electrons go into the L shell. This shell is completely filled in the case of neon (atomic number 10). Thus, neon has 2 electrons in the K shell and 8 in the L shell.

The arrangement of electrons in the different shells is known as the *electronic configuration* of the elements. The electronic configurations of the first eighteen elements are given in Table 3.5.

TABLE 3.5  
Electronic configurations of the first 18 elements

Element	At. No.	No. of electrons	Electronic configuration				
			K	L	M	N	O
Hydrogen ...	1	1	1	—	—	—	—
Helium ...	2	2	2	—	—	—	—
Lithium ...	3	3	2	1	—	—	—
Beryllium ...	4	4	2	2	—	—	—
Boron ...	5	5	2	3	—	—	—
Carbon ...	6	6	2	4	—	—	—
Nitrogen ...	7	7	2	5	—	—	—
Oxygen ...	8	8	2	6	—	—	—
Fluorine ...	9	9	2	7	—	—	—
Neon ...	10	10	2	8	—	—	—
Sodium ...	11	11	2	8	1	—	—
Magnesium ...	12	12	2	8	2	—	—
Aluminium ...	13	13	2	8	3	—	—
Silicon ...	14	14	2	8	4	—	—
Phosphorus ...	15	15	2	8	5	—	—
Sulphur ...	16	16	2	8	6	—	—
Chlorine ...	17	17	2	8	7	—	—
Argon ...	18	18	2	8	8	—	—



## QUESTIONS

1. What evidence led to the discovery of positively charged particles in the atom?
2. In a gas discharge tube, what is the purpose of using a perforated cathode?
3. What is meant by radioactivity? Name any two radioactive elements that you know of.
4. Briefly describe the nuclear model of the atom. What evidence led to this model?
5. The atomic number of helium is 2 but its mass number is 4. What is the composition of the nucleus?
6. Fill in the blanks in the following table :

Element	Atomic number	Protons	Electrons	Neutrons	Mass number
Magnesium	12	12	12	12	24
Nitrogen	7	7	7	7	14
Chlorine	17	17	17	18	35
Aluminium	13	13	13	14	27
Phosphorus	15	15	15	16	31
Calcium	20	20	20	20	40

7. Could neutral atoms of the same element have

- (a) different number of protons?
- (b) different number of neutrons?
- (c) different number of electrons?
- (d) different mass numbers?

Explain your answers.

8. Could neutral atoms of different elements have the

- (a) same number of protons?
- (b) same number of neutrons?
- (c) same number of electrons?
- (d) same mass number?

Explain your answers.

9. What information is conveyed by the symbol  ${}_7\text{N}^{15}$ ?
10. Oxygen has atomic number 8 and mass number 16. What is the number of neutrons in its nucleus?
11. Write down the electronic configurations for the following elements. The atomic number of the element is given in the brackets.  
Helium (2), Carbon (6), Fluorine (9) and Magnesium (12).
12. Isotopes of an element differ in the
  - (a) number of protons
  - (b) number of neutrons
  - (c) number of electrons.
13. Match the appropriate phrases from group B with the words in group A.

A

B

- |                       |                                                                                    |
|-----------------------|------------------------------------------------------------------------------------|
| 1. $\alpha$ -Particle | (a) Particle obtained when hydrogen atom loses an electron                         |
| 2. Proton             | (b) Particle having a mass equal to that of a hydrogen atom but carrying no charge |
| 3. Neutron            | (c) Particle obtained when helium atom loses two electrons                         |
| 4. Atomic number      | (d) Atoms with the same atomic number but different mass numbers.                  |
| 5. Isotopes           | (e) Small region in the atom containing protons and neutrons                       |
| 6. Nucleus            | (f) Net positive charge on the nucleus of an atom                                  |

14. State whether the following statements are TRUE or FALSE
  - (a) A proton has a mass equal to that of a hydrogen atom and carries a negative charge.
  - (b) A neutron carries no charge but has a mass equal to that of an electron
  - (c) The number of positive charges in an atom of an element is known as its atomic number.
  - (d) The L-shell can accommodate only 2 electrons.
  - (e) Positively charged particles are formed when electrons are lost by neutral atoms.
  - (f) The fractional atomic weights of elements are due to the existence of isotopes.
  - (g) The nucleus contains the total negative charge of the atom.

5. Fill in the blanks with suitable words :

- (a) K shell can accommodate.....electrons.
- (b) Oxygen with atomic number 8 and mass number 16 has.....protons ;  
.....electrons and.....neutrons.
- (c) The size of the nucleus is .....than that of the atom.
- (d) The mass of an electron is.....of that of a proton.
- (e) In a neutral atom the number of.....is equal to the number of.....
- (f) Beta rays contain , .....charged particles.
- (g) The region of an atom in which protons are found is called.....
- (h) The nucleus of an atom contains.....and.....

## 4

# HOW ATOMS COMBINE

Atoms consist of different particles. The arrangement of these particles in atoms has been discussed in the previous chapter. Atoms combine to form molecules. The different ways in which they do so are dealt with in this chapter.

## 4.1 Combination between atoms

Many attempts have been made to explain the chemical combination between atoms to form molecules. But a simple explanation was possible only after knowing how electrons are arranged in atoms.

The elements which have practically no tendency to react chemically are the noble gases. The electronic configurations of these gases are given in Table 4.1.

TABLE 4.1  
Electronic configurations of the noble gases

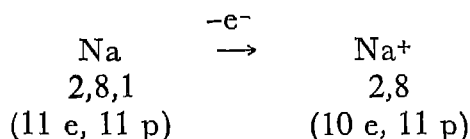
Noble gas	At. No.	Configuration
Helium	2	2
Neon	10	2,8
Argon	18	2,8,8
Krypton	36	2,8,18,8
Xenon	54	2,8,18,18,8
Radon	86	2,8,18,32,18,8

Except helium, each of these elements has an arrangement of eight electrons in the outermost shell (which is also called the valence shell). From this we may guess that this type of configuration represents marked stability. The atoms of other elements have a configuration different

from that of the noble gases. For example, the sodium atom has the configuration 2,8,1. It has one electron more than the neon atom. The chlorine atom with a configuration 2,8,7, has one electron less than an argon atom. Whenever a chemical reaction takes place these atoms tend to acquire a stable configuration of the noble gas type. Let us consider some of the possible ways of attaining this.

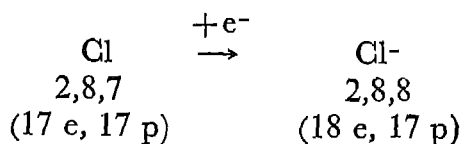
## 4.2 Ionic bond

A sodium atom can acquire the stable configuration of a noble gas either by losing one electron or gaining seven. It is much easier for the sodium atom to lose the single valence electron rather than gain seven. By doing so it acquires the neon-like configuration, and forms a positive sodium ion,  $\text{Na}^+$ .



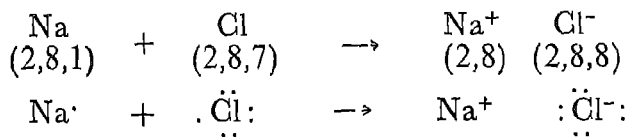
Due to the loss of an electron there now remain only 10 electrons, whereas the nucleus contains 11 protons. There is thus an excess positive charge of one unit.

When a chlorine atom gains an electron it acquires an argon-like configuration and forms a negative chloride ion,  $\text{Cl}^-$ .



Due to the gain of an electron there are now 18 electrons while the nucleus has 17 protons. Thus, there is an excess negative charge of one unit.

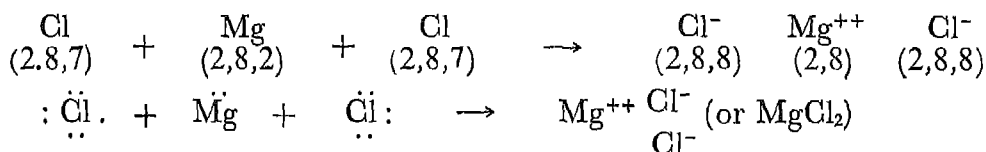
Now when sodium and chlorine react, each sodium atom transfers an electron to a chlorine atom. As a result a positive sodium ion,  $\text{Na}^+$ , and a negative chloride ion,  $\text{Cl}^-$ , are formed. These two ions are held together by the electrostatic attraction of their opposite charges, forming sodium chloride.



The electrostatic force which holds the two ions together is known as an '*ionic bond*'. The compound thus formed is called an *ionic compound*.

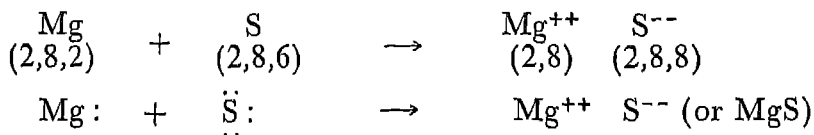
The single dot around sodium and the seven dots around chlorine represent the electrons in the outermost shell of a sodium and chlorine atom respectively.

Take the case of the formation of magnesium chloride. The magnesium atom has the configuration 2,8,2. It has two electrons more than the neon atom. When it combines with chlorine, it transfers these two electrons to two chlorine atoms, one to each. The magnesium ion,  $\text{Mg}^{++}$ , thus formed has a neon-like configuration and each of the chloride ions, an argon-like configuration. The formation of the ionic compound magnesium chloride may be represented thus:



The magnesium ion,  $\text{Mg}^{++}$ , has two positive charges. This is balanced by a total of two negative charges, one on each chloride ion. Thus, on the whole the compound is electrically neutral.

Another example of an ionic compound is magnesium sulphide.



In this case, the valence electrons of a magnesium atom are transferred to a sulphur atom. The magnesium ion has two positive charges

---

\* The notation used to represent the formation of molecules is called *electron dot representation*. It should not be supposed, however, that this type of notation represents the actual picture of the molecules. Nor are the electrons static. This notation is merely used for convenience and easy understanding. You will learn more about the nature of electrons and their distribution in atoms and molecules in higher classes.

and the sulphide ion,  $S^{--}$ , two negative charges. Again the compound as a whole is neutral.

Thus, the formation of ionic compounds involves the transfer of one or more electrons from the atom of one element to the atom of another. The atom losing the electrons becomes a positive ion and the atom receiving them becomes a negative ion.

### 4.3 Covalent bond

Let us consider the formation of the hydrogen molecule,  $H_2$ . Each hydrogen atom has one electron in its valence shell and needs one more electron to acquire a helium-like configuration. This can be achieved if the two atoms mutually share their two electrons. This electron pair is then simultaneously attracted by two hydrogen nuclei, which are consequently held together in the form of the molecule  $H_2$ . The shared pair of electrons belongs to both atoms. Each hydrogen atom can be considered to have acquired a helium-like configuration.



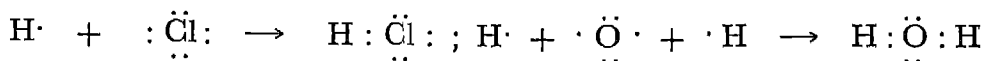
The bond thus formed by mutual sharing of a pair of electrons is a *covalent bond*. The molecule formed is a *covalent molecule*.

Let us consider another example, the formation of the chlorine molecule. Each chlorine atom needs one more electron to attain the argon-like configuration. This can be achieved if each chlorine atom shares one of its electrons with the other chlorine atom.

The combination of two chlorine atoms may be represented by the following equation.



The following equations show the formation of the molecules of hydrogen chloride and water.



Thus, the formation of covalent compounds involves the sharing of electrons between two atoms. Each atom contributes an electron to the

shared pair. A covalent bond is sometimes represented by means of a dash (—) between the bonded atoms. Thus, a molecule of hydrogen is represented as H—H, and a molecule of hydrogen chloride as H—Cl.

#### 4.4 Polarity of a covalent bond

In a hydrogen or chlorine molecule the pair of electrons forming the covalent bond is shared by identical atoms. This being so, the two atoms have an equal attraction for the electrons. Therefore, the sharing of electrons is also equal. On the other hand, in a molecule of hydrogen chloride the pair of electrons is shared by two dissimilar atoms. The chlorine atom has a greater attraction for electrons than hydrogen. As a result, the shared pair of electrons shifts nearer to the chlorine atom. This leads to unequal sharing of electrons between chlorine and hydrogen. The chlorine end of the molecule becomes slightly negative and the hydrogen end slightly positive. This is indicated as



The symbol  $\delta^+$  (delta plus) indicates that hydrogen has a slight positive charge and the symbol  $\delta^-$  (delta minus) indicates that chlorine has a slight negative charge in the molecule. Bonds formed by an unequal sharing of electrons are called *polar bonds*. In the water molecule oxygen forms two covalent bonds, one each with a hydrogen atom. It has greater attraction for electrons than hydrogen. Therefore, in the water molecule the oxygen end becomes slightly negative and the hydrogen end slightly positive. Thus, water is an example of a *polar* molecule. A convenient way of representing a polar molecule is indicated below.



We find that a number of covalent compounds are polar. ~~The~~ The polar nature affects the properties of the compounds to some extent. For instance, the behaviour of water as a solvent is mainly due to the polar nature of its molecules. You will learn about this in the succeeding chapters.

Molecules in which there is no net separation of positive and negative charges are called *nonpolar* molecules. Some examples of polar and nonpolar molecules are given below.



*Polar molecules*

Hydrogen chloride, HCl

Hydrogen bromide, HBr

Water, H<sub>2</sub>O

Ammonia, NH<sub>3</sub>

Hydrogen sulphide, H<sub>2</sub>S

*Nonpolar molecules*

Hydrogen, H<sub>2</sub>

Oxygen, O<sub>2</sub>

Chlorine, Cl<sub>2</sub>

Nitrogen, N<sub>2</sub>

Methane, CH<sub>4</sub>

Carbon tetrachloride, CCl<sub>4</sub>

QUESTIONS

1. In the formation of sodium chloride, how do sodium and chlorine atoms attain the noble gas-like configuration? Explain.
2. In the formation of hydrogen and chlorine molecules how do hydrogen and chlorine atoms acquire the configuration of a noble gas?
3. What is an 'ion'?  
How does an ion differ from an atom?
4. What are polar molecules?
5. Positive ions are formed from neutral atoms by the loss of
  - (a) neutrons
  - (b) protons
  - (c) electrons
6. Fill in the blanks in the following statements :
  - (a) Negative ions are formed by the.....of.....
  - (b) In the hydrogen chloride molecule the electrons are attracted more towards ..... than towards .....
  - (c) The atomic number of magnesium is 12. The magnesium ion, Mg<sup>++</sup>, has the configuration....
  - (d) The atomic number of chlorine is 17. The chloride ion, Cl<sup>-</sup>, has the configuration.....
  - (e) All noble gases except helium have ..... electrons in their outermost shell.
  - (f) Covalent molecules are formed by.... of electrons
  - (g) Ionic bond results when ions of opposite.....attract each other,

7. Write TRUE or FALSE against the following statements .

- (a) An atom of sulphur gains 2 electrons to form a sulphide ion,  $S^{2-}$ .
- (b) Magnesium chloride is a covalent compound.
- (c) In the water molecule, hydrogen has greater attraction for electrons than oxygen.
- (d) Carbon tetrabromide,  $CBr_4$ , is a polar molecule.
- (e) In an iodine molecule,  $I_2$ , the two atoms have equal attraction for the electrons.
- (f) Water is a nonpolar molecule.
- (g) Magnesium sulphide is formed by transfer of two electrons from a magnesium atom to a sulphur atom.
- (h) The electronic configuration of sodium ion is identical to that of argon.

9. Match the following :

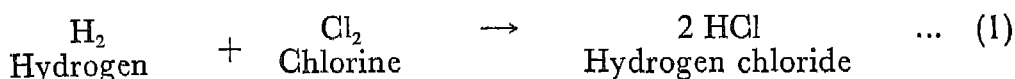
- |                               |                         |
|-------------------------------|-------------------------|
| 1. Polar covalent molecule    | a. Magnesium chloride   |
| 2. Nonpolar covalent molecule | b. Water                |
| 3. Ionic compound             | c. Carbon tetrachloride |

## CHEMICAL EQUATIONS

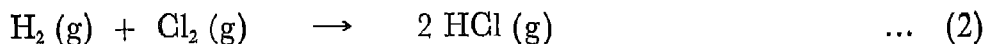
Elements and compounds are represented by symbols and formulae respectively. Similarly, reactions are represented by chemical equations. We shall study in this chapter about the significance and meaning of chemical equations.

### 5.1 Meaning of an equation

The reaction between hydrogen and chlorine resulting in the formation of hydrogen chloride may be represented by the following equation.



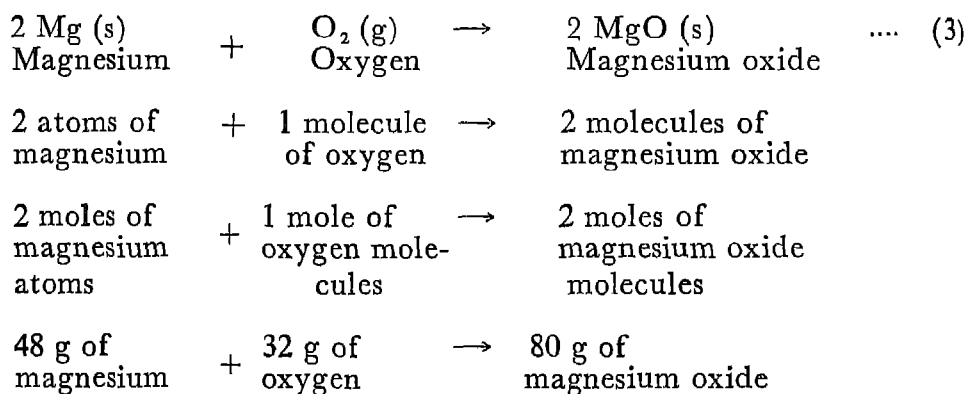
The above equation, however, does not indicate whether a particular reactant or product is present in a solid, liquid or gaseous state or in aqueous solution. For this purpose notations (s), (l), (g) and (aq) are used to represent a solid, liquid, gaseous and aqueous solution of the reactants or products. Accordingly the above equation may be written as



There are several ways in which we may interpret this equation. Qualitatively, the equation reads: gaseous hydrogen reacts with gaseous chlorine to give gaseous hydrogen chloride. Quantitatively: 1 molecule of hydrogen reacts with 1 molecule of chlorine to yield 2 molecules of hydrogen chloride. In actual practice, however, finite quantities of substances are studied; but even the smallest amount that can be easily handled consists of a very large number of molecules. It is therefore more convenient to use the *mole* ( $6.0 \times 10^{23}$  molecules) as the unit. Hence, the above equation reads: 1 mole of hydrogen reacts with 1 mole of chlorine to yield 2 moles of hydrogen chloride. In other words,  $6.0 \times 10^{23}$  molecules of hydrogen react with  $6.0 \times 10^{23}$  molecules of chlorine to give  $2 \times (6.0 \times 10^{23})$  molecules of hydrogen chloride. We may give the weights corresponding to the moles: 2 g of hydrogen react with 71 g of chlorine to yield 73 g of hydrogen chloride.

Usually it is more convenient to measure the volumes of gases rather than their masses. *Experiments have shown that one mole of every gas at the same temperature and pressure occupies the same volume (Avogadro's hypothesis).* Therefore, we can conclude from the above equation that 1 volume of hydrogen reacts with 1 volume of chlorine to yield 2 volumes of hydrogen chloride (provided all gases are measured at the same temperature and pressure).

In the case of solids and liquids masses rather than volumes, are usually measured. Consider the equation for the reaction of the burning of magnesium in oxygen.



Note that in each case the number of atoms and their total masses are conserved.

Equations must be used with caution. Because an equation can be written it does not follow that the reaction is possible. Even if the reaction does occur the equation gives no information about the conditions of temperature, pressure, etc, which may be necessary. It gives no information about the rate at which the reaction proceeds. It also gives no indication about the way in which the reactants are able to interact (mechanism). All one can say from an equation is the quantities of reactants and of products involved and their states of aggregation.

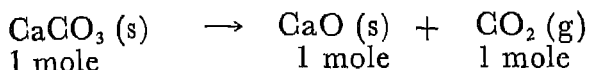
## 5.2 Problems based on equations

Chemical equations relate the masses and/or volumes of the reactants and products. These quantitative relations are termed *Stoichiometric relations*.

*Example: 1* How many grams of carbon dioxide gas will be produced by the decomposition of 10 g of calcium carbonate?

(Atomic weights: Ca = 40.0; C = 12.0; O = 16.0)

The chemical equation for the above reaction is



$$\begin{aligned} \text{Weight of 1 mole of calcium carbonate} &= \text{Ca} + \text{C} + 3 \text{ O} \\ &= 40 + 12 + 3 \times 16 \\ &= 40 + 12 + 48 \\ &= 100 \text{ g} \end{aligned}$$

Let us first determine the number of moles of calcium carbonate,  $\text{CaCO}_3$ , present in 10 g of this compound.

$$\begin{aligned} \text{No. of moles of CaCO}_3 &= \frac{\text{Weight of calcium carbonate}}{\text{Weight of 1 mole of calcium carbonate}} \\ &= \frac{10 \text{ g}}{100 \text{ g/mole}} \\ &= 0.1 \text{ mole} \\ 10 \text{ g of CaCO}_3 &= 0.1 \text{ mole of CaCO}_3 \end{aligned}$$

According to the above equation, 1 mole of  $\text{CaCO}_3$ , on decomposition, yields 1 mole of carbon dioxide,

$$\begin{aligned} \therefore 0.1 \text{ mole of CaCO}_3 \text{ will yield } & \frac{1 \times 0.1}{1} \\ &= 0.1 \text{ mole of carbon dioxide} \end{aligned}$$

But we know that 1 mole of carbon dioxide corresponds to 44 g of carbon dioxide.

$$\begin{aligned} \therefore \text{Weight of 0.1 mole of carbon dioxide} &= 0.1 \times 44 \\ &= 4.4 \text{ g} \end{aligned}$$

The weight of carbon dioxide produced from the decomposition of 10 g of  $\text{CaCO}_3$  is 4.4 g.

*Example: 2* 1.22 g of hydrated barium chloride when heated gives 1.04 g anhydrous barium chloride ( $\text{BaCl}_2$ ). Calculate the moles of water per mole of barium chloride. What is the formula of hydrated barium chloride?

(Atomic weights : Ba = 137, Cl = 35.5, H = 1, O = 16).

Wt. of hydrated barium chloride = 1.22 g

Wt. of anhydrous barium chloride = 1.04 g

∴ Wt. of water lost = 0.18 g

No. of moles of water in 0.18 g of water

$$= 0.18/18 \text{ mole}$$

$$= 0.01 \text{ mole}$$

Wt. of 1 mole of anhydrous barium chloride ( $\text{BaCl}_2$ )

barium chloride = Ba + 2 Cl

$$= 137 + 2 \times 35.5$$

$$= 208 \text{ g}$$

∴ No. of moles of anhydrous barium chloride

$$\text{in 1.04 g of the salt} = 1.04/208 \text{ mole}$$

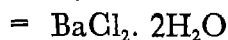
$$= 0.005$$

0.005 mole of  $\text{BaCl}_2$  is associated with 0.01 mole  $\text{H}_2\text{O}$

∴ 1 mole of  $\text{BaCl}_2$  is associated with

$$0.01/0.005 = 2 \text{ moles of } \text{H}_2\text{O}$$

Hence, the formula of hydrated barium chloride

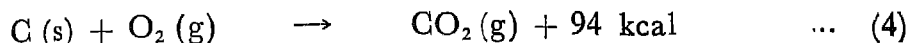


### 5.3 Equations showing the heat of reaction

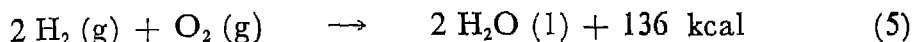
Nearly all chemical reactions involve some change in heat energy. There are various ways of indicating the quantity of heat that is liberated or absorbed. Heat is usually measured in calories. As the heat changes in chemical reactions are large, the unit used is kilocalorie (1 kcal = 1000 calories).

For example, consider the equation for the reaction of the burning of carbon in oxygen. It has been observed that 1 mole of carbon combines with 1 mole of oxygen to form 1 mole of carbon dioxide liberating 94 kcal of heat.

We express this fact as

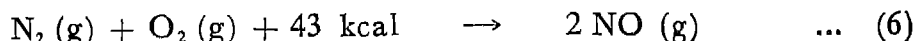


Consider when 2 moles of gaseous hydrogen react with 1 mole of gaseous oxygen to form 2 moles of liquid water, 136 kcal of heat are liberated. We express this fact as



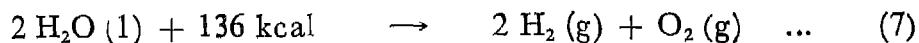
The amount of heat liberated in these reactions is shown on the right. Such reactions are called *exothermic reactions*.

When nitrogen combines with oxygen to form nitrogen monoxide, NO, heat is absorbed. This is indicated by showing the heat on the left as shown in the following equation.



The equation (6) indicates that 1 mole of gaseous nitrogen and 1 mole of gaseous oxygen absorb 43 kcal of heat and react to form 2 moles of gaseous nitrogen monoxide, NO. This is an example of an *endothermic reaction*.

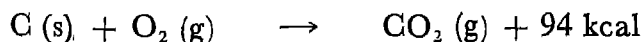
The decomposition of water to gaseous hydrogen and oxygen is another example of an endothermic reaction.



The equation (7) indicates that for the complete decomposition of 2 moles of liquid water into 2 moles of gaseous hydrogen and 1 mole of gaseous oxygen, 136 kilocalories of heat are absorbed. It may be noted that the same amount of heat is liberated in the reverse reaction as shown in equation (5).

*Example : 3* Calculate the amount of heat liberated when 12 kg of coal (carbon) are burnt completely.

The equation for the above reaction is



First determine the number of moles of carbon present in 12 kg *i.e.* 12,000 g of carbon.

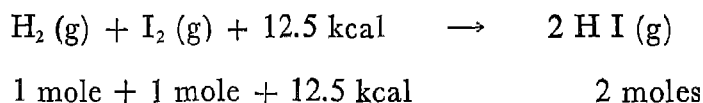
$$\text{No. of moles of carbon} = \frac{12000 \text{ g}}{12 \text{ g/mole}} = 1000 \text{ moles}$$

According to the equation 1 mole of carbon on complete combustion liberates 94 kcal. Hence 1000 moles of carbon would yield 94,000 kcal of heat.

*Example : 4* How much heat is required for the production of 12.8 g of hydrogen iodide from hydrogen and iodine?

(Atomic weight of I = 127, H = 1).

The equation for this reaction is



No. of moles of hydrogen iodide present in 12.8 g of HI  
 $= 12.8 / 128 = 0.1 \text{ mole}$

2 Moles of hydrogen iodide, HI, require 12.5 kcal

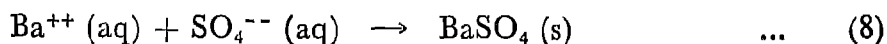
$\therefore$  0.1 mole of hydrogen iodide will require

$$\frac{12.5 \times 0.1}{2} = 0.63 \text{ kcal}$$

0.63 kcal of heat should be supplied to produce 12.8 g of hydrogen iodide from hydrogen and iodine

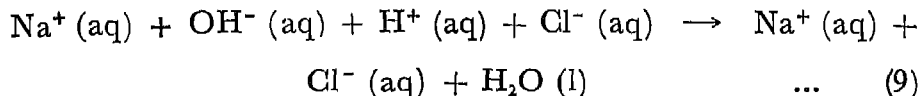
## 5.4 Ionic equations

A number of chemical reactions take place readily in aqueous solution. Many of these reactions involve ions and are, therefore, called *ionic reactions*. It is observed that a white precipitate of barium sulphate is obtained when an aqueous solution of barium chloride is mixed with an aqueous solution of a sulphate. It does not matter what barium salt (barium chloride or nitrate) is used or what sulphate (sodium or potassium or ammonium sulphate) is used. The other ions, namely,  $\text{Na}^+$ ,  $\text{K}^+$ , or  $\text{NH}_4^+$  and  $\text{Cl}^-$  or  $\text{NO}_3^-$ , remain in solution as before, that is, they do not take part in the reaction. Such ions are called "spectator ions". The actual reaction is between barium ions and sulphate ions.





We are familiar with the reaction between sodium hydroxide, NaOH, and hydrochloric acid, HCl, leading to the formation of sodium chloride and water. This is known as *neutralization*. Sodium hydroxide, hydrochloric acid and sodium chloride are completely ionized in solution. The reaction is represented as follows :



We find that the sodium ion,  $\text{Na}^+$ , and the chloride ion,  $\text{Cl}^-$ , do not take part in the reaction. To neutralize a base, it does not matter what acid is used. Likewise an acid can be neutralized with any base. In fact, the actual reaction that takes place during neutralization is the combination of  $\text{H}^+$  (aq) ions and  $\text{OH}^-$  (aq) ions to form water.

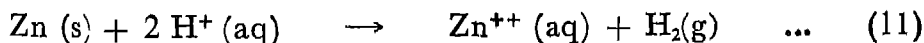


The equation (10) tells us that 1 mole of hydrogen ions,  $\text{H}^+$  (1 g) reacts with 1 mole of hydroxyl ions,  $\text{OH}^-$  (17 g) giving 1 mole (18 g) of liquid water. We find that the mass is conserved, as in all cases.

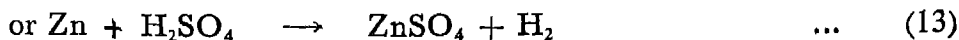
The positive charge on the hydrogen ion balances the negative charge on the hydroxyl ion leaving a net zero charge on liquid water. We have thus not only a mass balance, but also a charge balance.

The familiar reaction between zinc and dilute acid is also an ionic reaction. The same reaction occurs whether the acid is hydrochloric or sulphuric. Here the metal dissolves and hydrogen is evolved. Zinc goes into solution as zinc ions,  $\text{Zn}^{++}$  (aq), and the hydrogen ions,  $\text{H}^+$  (aq), in the acid are converted to hydrogen gas.

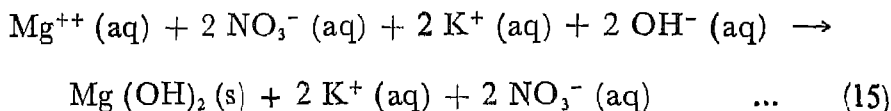
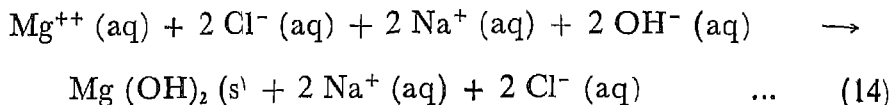
Hence, the equation is



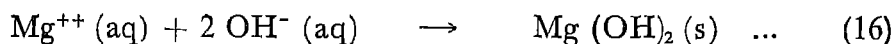
This is a truer representation than



Consider the following reactions :



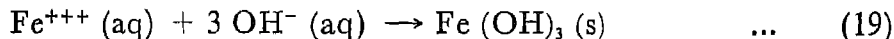
We find from the equations (14) and (15) that magnesium hydroxide  $\text{Mg}(\text{OH})_2$ , is precipitated in each case. It does not matter what ions other than  $\text{Mg}^{++}(\text{aq})$  and  $\text{OH}^{-}(\text{aq})$  ions are present in the aqueous solution. They do not take part in the reaction. Hence the single observable change for both the reactions can be represented thus :



Similarly, lead ions react with iodide ions to form a yellow precipitate of lead iodide,  $\text{PbI}_2$ .



Iron (III) ions react with hydroxyl ions to form a reddish-brown precipitate of iron (III) hydroxide,  $\text{Fe}(\text{OH})_3$ .



The examples mentioned above suggest that in ionic reactions only the participating ions and the products need be indicated. While writing these equations it is necessary that the number of atoms as well as the charges are balanced.

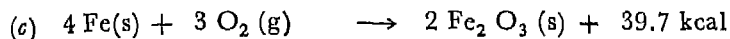
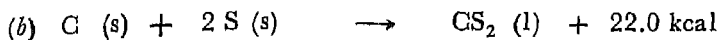
In writing ionic equations, a convention is adopted. For example, if an insoluble substance like barium sulphate, lead iodide or iron (III) hydroxide is produced in a reaction, that substance is shown in an electrically neutral, molecular form, though it is an ionic solid.

## QUESTIONS

1. State all the information provided by the equation



2. Explain the meaning of the following equations.



3. What weight of zinc oxide will be formed when 32.7 g of zinc react with oxygen?  
(Zn = 65.4, O=16.0)

4. Calculate the number of moles of mercuric oxide, HgO, needed to produce 8 g of oxygen.  
(Hg = 200.6, O=16.0)

5. Calculate the number of moles of oxygen that react with 60 g of carbon to form carbon dioxide  
(C = 12.0 O = 16.0)

6. What are exothermic reactions? Give examples.

7. What are endothermic reactions? Give examples.

8. Classify the following as endothermic or exothermic processes.

(a) Decomposition of water

(b) Dissolution of ammonium chloride in water

(c) Explosion of a bomb

(d) Formation of slaked lime from quicklime

9. How much heat is required for the decomposition of 18 g of water into gaseous hydrogen and oxygen?



10. When 63.54 g of copper react with sulphur to form copper (II) sulphide, 11.6 kcal of heat are liberated. If 19.071 g of copper completely react to form CuS, calculate the quantity of heat liberated.



11. Calculate the percentage gain in weight when 1.0 g of magnesium burns completely in oxygen to form 1.65 g of magnesium oxide.  
(Mg = 24.3, O = 16.0)

12. 2.495 g of copper (II) sulphate pentahydrate, when heated, gives anhydrous copper sulphate. Calculate the weight of the anhydrous copper (II) sulphate.  
(Cu = 63.5, S = 32.0, O = 16.0, H = 1.0)

13. What are ionic equations?

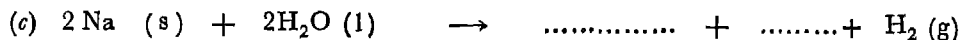
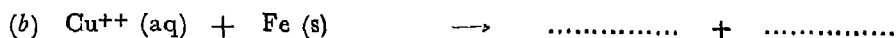
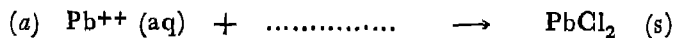
14. Write the net ionic equations for the following :

(a) Black copper (II) sulphide is precipitated when hydrogen sulphide, is passed through a solution of copper (II) sulphate.

(b) Bromine displaces iodine from potassium iodide solution.

(c) Magnesium dissolves in dilute sulphuric acid giving hydrogen gas and magnesium sulphate.

15. Fill in the following blanks.



## HYDROGEN, OXYGEN AND WATER

Hydrogen and oxygen are two important elements, finding a wide use in industry. Oxygen is absolutely essential to life. Water, by far the most widely occurring compound of hydrogen and oxygen, is indispensable to life and industry. In this chapter we deal with some important aspects of these elements and their important compounds.

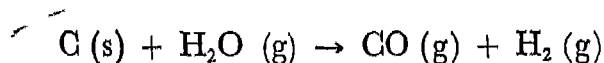
### 6.1 Hydrogen

Hydrogen is used in the manufacture of a large number of industrially useful compounds. Under appropriate conditions hydrogen combines with nitrogen to form ammonia, used in the manufacture of nitric acid and fertilizers. Hydrogen is also used in the hardening of oils and in the oxy-hydrogen welding torch.

#### *Preparation of hydrogen*

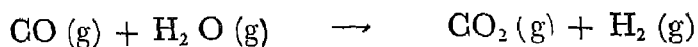
Hydrogen is prepared in the laboratory by the action of dilute acids on metals like tin, magnesium, zinc, etc. This method, however, is not suitable for the large scale preparation of the gas because of the high cost of the materials involved. For industrial preparation cheaper raw materials have to be used. A few industrial methods of preparing hydrogen are mentioned below :

(i) *By the action of steam on coke :* When steam is passed over red hot coke (which is largely carbon), a mixture of carbon monoxide and hydrogen is obtained,



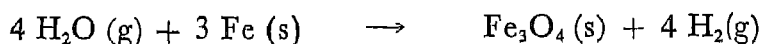
This mixture is known as *water gas* and is used either as a fuel, or in the manufacture of hydrogen itself, as described below.

A mixture of water gas and steam is passed over a catalyst, iron (III) oxide (mixed with chromium (III) oxide) at 450 to 500° C. The carbon monoxide is oxidized by steam to carbon dioxide and steam is reduced to hydrogen.

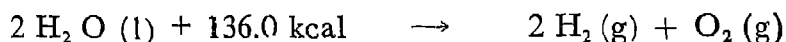


The carbon dioxide formed is removed by dissolving it, under pressure, in water. The trace of carbon monoxide still remaining is removed by washing the gas, under pressure, with ammoniacal copper (I) formate.

(ii) *By the action of steam on iron* : When steam is passed over red hot iron (600-850°C) hydrogen is obtained.



(iii) *By the electrolysis of water* : This method is used wherever electricity is cheap. Oxygen is an important by-product.



### *Properties of hydrogen*

Hydrogen is a colourless, tasteless and odourless gas insoluble in water. It is the lightest of all elements and has the simplest atomic structure. The nucleus of its most abundant isotope, having a mass of one unit, consists of a single proton. The gas exists in the stable form as a diatomic molecule,  $\text{H}_2$ .

In its chemical reactions with non-metals, hydrogen usually attains a stable helium-like configuration by sharing its electron with that of another atom, forming a single covalent bond.

At room temperature, hardly any reaction takes place between hydrogen and oxygen. But if a mixture of the two gases is ignited by a spark, or if a flame is brought in contact with the mixture, the gases react with explosion to form water.



Hydrogen combines with chlorine in diffuse daylight forming hydrogen chloride.



The reaction occurs violently in sunlight or other bright light. Bromine and iodine react with hydrogen only when heated.

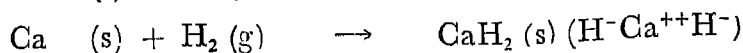
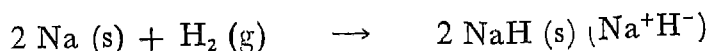
When hydrogen is bubbled through molten sulphur, a foul smelling gas, hydrogen sulphide, is formed.



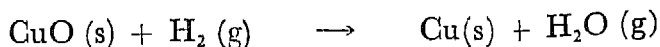
Hydrogen reacts with nitrogen at higher temperatures and under the influence of a catalyst, and forms ammonia.



It reacts with metals such as sodium, potassium, calcium, etc., to form the corresponding hydrides. It is present in these hydrides as a hydride ion,  $\text{H}^-$ .



If hydrogen is passed over heated copper (II) oxide, the latter is reduced to metallic copper. Hydrogen is oxidized to water.



## 6.2 Oxygen

Oxygen is essential to life. To those whose lungs have been impaired, an oxygen-rich atmosphere (the oxygen tent, containing 40-60% oxygen) has often proved a means of saving life. At high altitudes there is not enough oxygen to support life. The mountain climber, therefore, depends on oxygen supply through the 'mask'

The burning of coal, oil, petrol, wood—in fact, all combustion—depends on oxygen. Combustion is more vigorous in pure oxygen than in ordinary air. Rockets carry liquid oxygen for the burning of their fuel.

Oxygen is extensively used in the oxy-acetylene torches for the cutting and welding of metals and alloys.

### *Industrial preparation of oxygen*

Air is the common source of oxygen required for industry. Air is liquefied by cooling and compressing. The liquid contains mostly liquid oxygen (b.p.  $-183^\circ\text{C}$ ) and liquid nitrogen ( $-196^\circ\text{C}$ ). When heated,

liquid nitrogen boils off first, leaving oxygen behind. This kind of separation, which makes use of the difference in boiling points of two liquids, is known as *fractional distillation*.

Oxygen for industrial purposes is sometimes obtained by the electrolysis of water

### *Properties of oxygen*

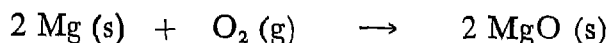
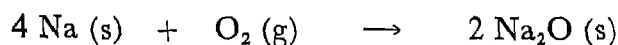
Oxygen is a colourless and odourless gas, slightly soluble in water. It exists in two molecular forms. The most common and stable form is the covalent diatomic molecule,  $O_2$ . The other form, known as ozone, has the formula  $O_3$ . When an element exists in more than one form, in the same physical state as, for example, oxygen and ozone, the phenomenon is known as *allotropy*. The different forms are called *allotropes*. Elements like carbon, sulphur and phosphorus also exist in several allotropic forms.

The oxygen atom has six valence electrons. A molecule of oxygen is formed by the sharing of two pairs of electrons. The structure of the molecule may be represented as



However, a number of its properties indicate that the above structure is not quite correct and that the structure of oxygen is more complex.

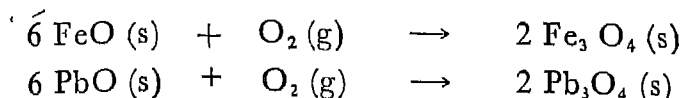
Under suitable conditions, oxygen reacts with most elements. With metals it forms basic oxides, in which it is largely present as the oxide ion,  $O^{--}$ .



It reacts with non-metals to form covalent molecules, most of which show acidic properties. For example, sulphur dioxide reacts with water to give an acidic solution (sulphurous acid,  $H_2SO_3$ ). However, some covalent oxides like carbon monoxide and water are neutral.



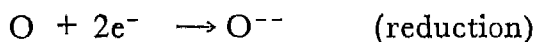
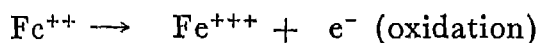
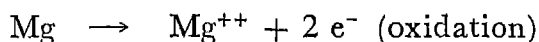
Oxygen also reacts with a number of compounds. For example, it combines with certain oxides to produce oxides which contain a higher percentage of oxygen.



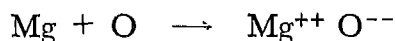
### 6.3 Oxidation and reduction

Magnesium oxide, MgO, is an ionic solid. Each magnesium ion has a net positive charge of two units and an oxide ion two units of negative charge. In the reaction between magnesium and oxygen to form magnesium oxide, a magnesium atom loses two electrons to an oxygen atom and forms the magnesium ion,  $\text{Mg}^{++}$ . The oxygen atom accepts two electrons and forms an oxide ion,  $\text{O}^{--}$ . Magnesium is said to be *oxidized* and oxygen *reduced*.

Oxidation is a process in which an atom or ion loses electrons. Reduction is a process in which an atom or ion gains electrons.



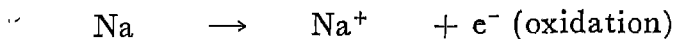
Reactions in which electrons are transferred from one atom to another are known as *oxidation-reduction reactions* or *redox reactions*. For example,



In this reaction magnesium is the *reducing agent*, since it has reduced oxygen to the oxide ion. On the other hand, oxygen is the *oxidizing agent*. It has oxidized magnesium to the magnesium ion. Therefore, *a reducing agent loses electrons and gets oxidized. An oxidizing agent gains electrons and gets reduced.* Oxidation and reduction occur simultaneously.

The formation of sodium chloride from sodium and chlorine is also a redox reaction.

Sodium atom is oxidized to a sodium ion



and a chlorine atom is reduced to a chloride ion



Sodium is the reducing agent and chlorine, the oxidizing agent.

## 6.4 Ozone

As stated already ozone,  $\text{O}_3$ , is an allotrope of oxygen. It is formed in small quantities in the atmosphere whenever lightning strikes. The fishy odour around high voltage electrical machinery is due to this gas.

Ozone finds use as a bleaching agent for oils, waxes, starch and flour. It is sometimes used to purify air and to sterilize water.

### *Preparation of ozone*

Ozone is prepared by passing a silent electric discharge\* through air or oxygen in a special type of apparatus called *ozoniser*. The *Siemen's*

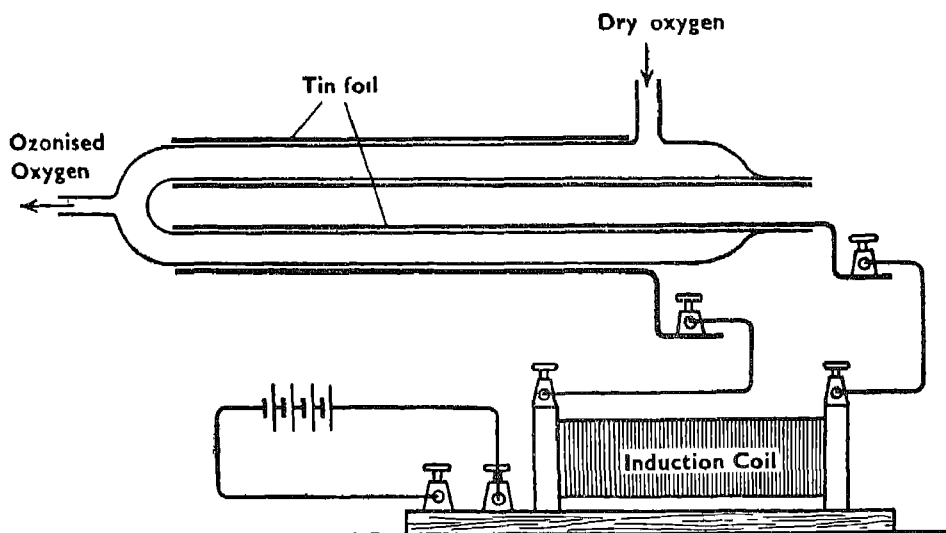


Fig. 6.1 Siemens, Ozoniser

\* A silent electric discharge differs from an ordinary spark discharge by the fact that the heat generated is considerably less in the former case. The silent discharge is obtained by inserting insulating materials in the space between the electrodes through which the discharge passes. The discharge then occurring through a wide surface prevents any local rise of temperature, and would lead to the decomposition of ozone.

*ozoniser* (Fig.6.1) consists of two coaxial glass tubes, the outside of the outer tube and the inside of inner tube being covered with tin foil to form the electrodes. The tin foils are connected to the terminals of an induction coil.

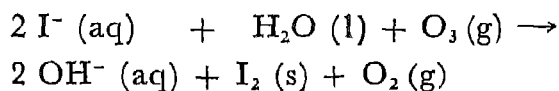
Pure dry oxygen is passed in a slow stream through the space between the two tubes. When the discharge takes place, some of the oxygen is converted to ozone. The formation of the gas is detected by its smell and by its action on starch-iodide paper which is turned blue.

### *Properties of ozone*

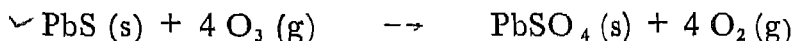
Ozone is a faintly bluish gas with a characteristic fishy odour (its name in greek means—to smell). It is slightly soluble in water but more so in carbon tetrachloride and benzene.

Ozone decomposes slowly at low temperatures, but almost at once at higher temperatures. It reacts with other substances more vigorously than oxygen does, and so is a stronger oxidizing agent. Some of its oxidizing reactions are mentioned below :

(i) Potassium iodide is oxidized to iodine



(ii) Lead sulphide (black) is oxidized to lead sulphate (white)



In all the reactions, ozone seems to break down into an oxygen molecule and an oxygen atom.



The atomic oxygen is the oxidizing agent.

## **6.5 Water**

Water plays a vital role in a large number of processes upon which life depends. It is widely distributed in nature in the liquid, solid and vapour states. As liquid it fills the oceans, lakes and rivers and covers nearly 5/7 of the earth's surface. In the form of ice it is present in large quantities in the polar regions and on the top of high mountains. Large

quantities of water vapour are present in the atmosphere. Water is also present in all living plants and animals and in many rocks and minerals.

Water occurs in nature as rain, river, spring, lake and sea water. Natural waters contain various impurities. The nature and amounts of these impurities depend upon the source. The only source of natural water which is relatively pure is rain water. The impurities in natural waters are of two kinds: (i) suspended—mineral and organic; and (ii) dissolved—solids and gases. Sea water is the most impure of natural waters. It contains about 3.5 per cent by weight of dissolved salts. A large portion of these is sodium chloride.

#### *Water for drinking purposes*

Water, as it occurs in nature, is not normally fit for drinking purposes. It contains many impurities and harmful bacteria. For drinking purposes it should be freed from these harmful ingredients. However, completely pure water has a rather insipid taste. Minute traces of dissolved salts (for instance, the bicarbonates) impart a refreshing taste to water and are not injurious to health.

In spite of extremely large quantities of water being available on the earth, the problem of obtaining drinking water is an acute one. In our country large areas do not have a regular supply of drinking water. In many big cities of the world the supply of drinking water poses problems to the authorities. Attempts are being made in some countries to use solar energy for the purification of water. Desalination of sea water (i.e. removal of dissolved salts) on a large scale is being attempted in a few countries. These two methods have large possibilities in our country too.

To be useful for drinking purposes (potable), water from natural sources, such as river or rain water, must be freed from suspended impurities and harmful bacteria. So it must be filtered and sterilized. Filtration is carried out through beds of gravel and sand which become covered with a layer of clay, algae, etc. This layer removes suspended matter and also many bacteria. A little aluminium sulphate may also be added to clear the water from finely suspended clay. Aluminium hydroxide is precipitated and carries down suspended particles and also some bacteria.

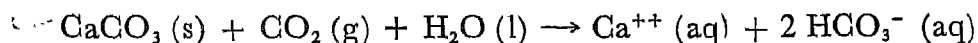
Sterilization of water is done by means of various chemicals. It is done by adding bleaching powder solution, or by passing chlorine gas, or

a mixture of chlorine and ammonia (chloramine process), or by bubbling ozonised air through filtered water. Sterilization is also done by irradiating water with ultraviolet rays. This method has the advantage that it does not impart any odour to water.

## 6.6 Hardness of water

Distilled water and rain water give lather readily with soap. Such water is called *soft water*. Water from the sea and some wells, instead of forming lather with soap, forms a scum or precipitate. Such water is called *hard water*. Hard water is saltish and sometimes bitter in taste.

Solutions containing bicarbonates, chlorides and sulphates of calcium and magnesium do not give lather readily with soap. But solutions containing other salts like sodium chloride do. It is evident that the ions like calcium and magnesium are responsible for the hardness of water. How do these ions get into water? The earth contains large amounts of calcium and magnesium salts. As the rain drops fall through air, they absorb carbon dioxide present in the air. The carbon dioxide dissolved in water (i.e. carbonic acid) then reacts with the carbonates of calcium and magnesium in the soil, forming a solution of their bicarbonates. Consequently water becomes hard.



In this way ground waters dissolve many such salts and the water becomes hard.

### *Disadvantages of using hard water*

When soap is treated with hard water, an insoluble precipitate is formed. Soap is sodium stearate (sodium salt of stearic acid,  $\text{C}_{17}\text{H}_{35}\text{COONa}$ ) and consists of positive sodium ions and negative stearate ions. The stearate ions of soap combine with the calcium or magnesium ions forming insoluble calcium or magnesium stearate. Lather is formed only after much soap is wasted, i.e., after the calcium and magnesium ions are removed as scum.

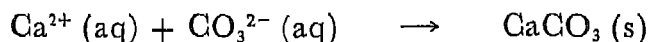
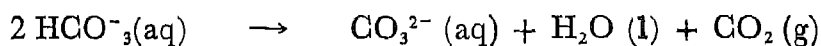
Hard water is also unsuitable for industrial use in boilers. When water containing calcium bicarbonate is heated, a deposit of insoluble calcium carbonate is formed on the inner walls of the boiler due to the decomposition of calcium bicarbonate. This deposit is known as *boiler scale*. In boilers even a thin deposit (scale) obstructs the transfer of heat

to water. More fuel is, therefore, required for converting water into steam. Sometimes boilers are put out of action by local overheating due to boiler scales.

### *Removal of hardness*

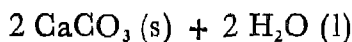
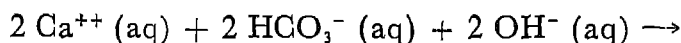
Hardness of water can be removed by eliminating the calcium and magnesium ions responsible for it. Some methods used for this purpose are outlined below :

a) *By boiling* : When water containing calcium and magnesium bicarbonates is heated, the salts decompose into insoluble calcium and magnesium carbonates. These settle down and can be filtered off.



This kind of hardness which is due to the bicarbonates of calcium and magnesium and which can be removed by mere boiling is called *temporary hardness*. When water containing chlorides and sulphates of calcium and magnesium is heated no insoluble material is precipitated. This kind of hardness is called *permanent hardness*.

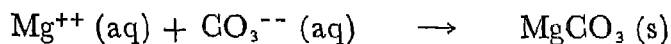
b) *By adding slaked lime* (calcium hydroxide,  $\text{Ca}(\text{OH})_2$ ) : When a calculated quantity of slaked lime is added to temporary hard water, the bicarbonates are precipitated as carbonates



Addition of slaked lime in excess of the requisite amount leaves free calcium ions in water, thereby making the water hard once again.

c) *By adding washing soda* : Permanent hardness cannot be removed by boiling or by adding lime.

It is removed by adding washing soda.



Washing soda also removes temporary hardness and so can be used to remove both kinds of hardness.

## 6.7 Composition of water

We have learnt earlier that the electrolysis of water gives oxygen and hydrogen. Hydrogen collects at the cathode and oxygen at the anode in the volume ratio of two parts of hydrogen to one part of oxygen (Fig. 6.2).

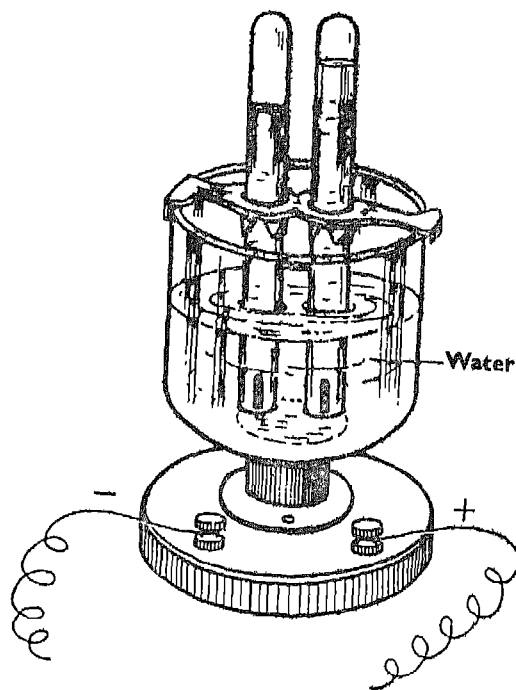


Fig. 6.2 Composition of water

Applying Avogadro's hypothesis to the above reaction, we find that:

2 molecules of hydrogen combine with 1 molecule of oxygen to form 2 molecules of water.

Since molecules of hydrogen and oxygen are diatomic, 2 molecules of hydrogen contain 4 hydrogen atoms and 1 molecule of oxygen contains 2 oxygen atoms. Therefore, in 2 molecules of water there must be 4 atoms of hydrogen and 2 atoms of oxygen. Thus, 1 molecule of water contains 2 atoms of hydrogen and 1 atom of oxygen. Hence, the formula of water is  $\text{H}_2\text{O}$ .

We know that under the influence of an electric spark, a mixture of hydrogen and oxygen explodes forming water. This reaction can also be used to determine the composition of water by volume.

For this purpose an *eudiometer* (Fig 6.3) is used. It consists of a graduated tube A with a uniform bore and connected to a mercury reservoir R. Two platinum wires are fused into the tube and are connected to an induction coil. A mixture containing known volumes of hydrogen and oxygen is introduced into the tube through the tap T. Hydrogen is taken in a large excess in the mixture. The volume of the mixture is noted after levelling the mercury in the tube and the reservoir.

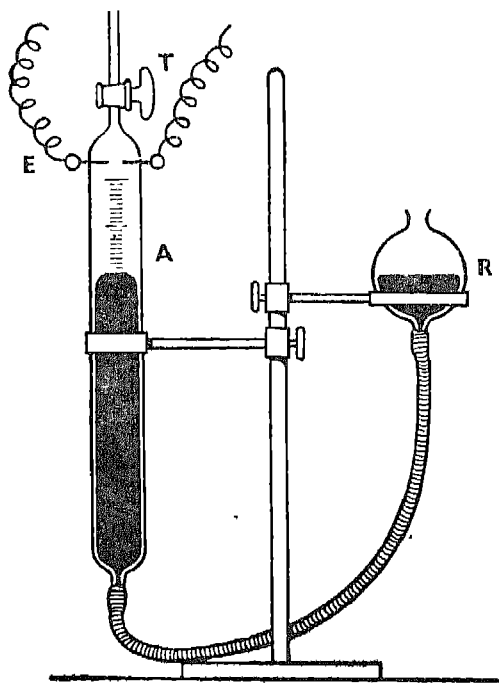


Fig. 6.3 Synthesis of water

The mixture is ignited by a spark. The two gases combine with an explosion to form water vapour. The water vapour condenses and the mercury rises in the tube. Only hydrogen remains in the tube after sparking (since it is taken in excess). The volume of the gas is measured after levelling the mercury once again. The volumes are then reduced to S.T.P and the volume composition of water is determined. The ratio of hydrogen to oxygen is found to be 2:1.



## 6.8 Properties of water

Pure water has no taste or smell. It freezes at  $0^{\circ}\text{C}$  and boils at  $100^{\circ}\text{C}$  at normal atmospheric pressure (760 mm mercury). Ice is less dense than water and hence floats on water. The density of most liquids decreases with rise of temperature. But the density of water increases from  $0^{\circ}\text{C}$  to  $4^{\circ}\text{C}$ , reaches a maximum value of 1 gram per ml at  $4^{\circ}\text{C}$  and then decreases as the temperature rises further.

### *Water as a solvent*

One of the most important properties of water is its capacity to dissolve a large number of substances. For this reason water is usually known as a *universal solvent*.

Let us consider how an ionic solid like sodium chloride dissolves in water (Fig. 6.4).

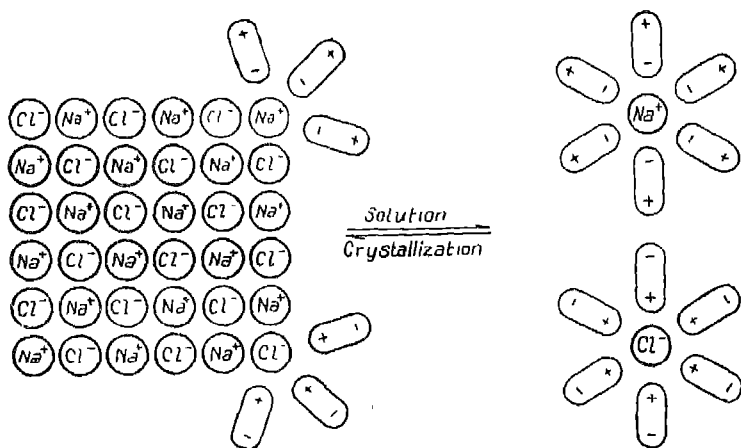
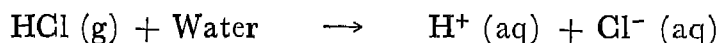


Fig. 6.4 Dissolution of sodium chloride in water

Sodium chloride consists of a regular arrangement of sodium ions,  $\text{Na}^+$ , and chloride ions,  $\text{Cl}^-$ , throughout the crystal. When sodium chloride is placed in water, the polar water molecules cluster around the sodium and chloride ions on the surface of the crystal. The positive ends of the water molecules are directed towards the chloride ions. Likewise the negative ends of water molecules are directed towards the sodium ions. This interaction between the polar water molecules and the ions leads to the *hydration* of these ions. During hydration a large amount of energy is released. This energy is used up in overcoming the electrostatic attrac-

tion between the oppositely charged ions in the crystal. As a result, the ions are torn apart and the solid dissolves in water.

Another important property of water as a solvent is its tendency to ionize acids, that is, to break down the molecules of an acid into ions. Hydrogen chloride is a covalent compound. When it is dissolved in water, the bond that holds the hydrogen and chlorine atoms together in the molecule is broken. The hydrogen comes off as a hydrogen ion,  $H^+$ , and the chlorine as a chloride ion,  $Cl^-$ . Both the ions are hydrated in solution.



### *Deliquescence and efflorescence*

Many substances absorb water from air. Common salt absorbs moisture when exposed to air. White, anhydrous copper (II) sulphate absorbs moisture and turns blue. Such substances are said to be *hygroscopic*. Lime and sulphuric acid are also hygroscopic.

Calcium chloride absorbs so much moisture from the atmosphere that it dissolves in the absorbed water. Such substances are said to be *deliquescent* (turning to liquid). Sodium hydroxide, potassium hydroxide and phosphorus (V) oxide are also deliquescent. All deliquescent substances are also hygroscopic. Hygroscopic and deliquescent substances are used as drying agents. Concentrated sulphuric acid and anhydrous calcium chloride are some of the drying agents which are usually employed.

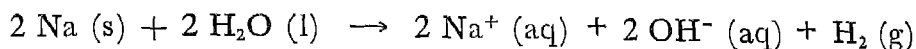
When a crystal of hydrated sodium thiosulphate,  $Na_2S_2O_3 \cdot 5 H_2O$ , is exposed to air, a white powder slowly forms on its surface. If left long enough, the entire crystal falls to a powder. This happens due to the loss of the water of hydration by the crystal.



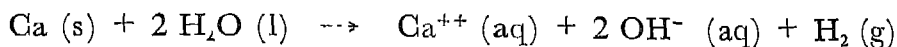
Such substances are said to be *efflorescent* and the phenomenon is called *efflorescence*. Sodium sulphate,  $Na_2SO_4 \cdot 10 H_2O$  and sodium carbonate,  $Na_2CO_3 \cdot 10 H_2O$ , are also efflorescent.

### *Action of water on metals*

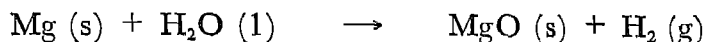
Metals like sodium and potassium react vigorously with water in the cold and liberate hydrogen.



Calcium, too, reacts with cold water, though less vigorously than sodium and potassium.



Cold water practically does not react with magnesium, but hot water does. Heated magnesium catches fire in steam and burns brightly.



Iron decomposes water extremely slowly during rusting. Rapid decomposition takes place when steam is passed over heated iron. Hydrogen and a bluish black iron oxide called *magnetic oxide of iron* are formed.

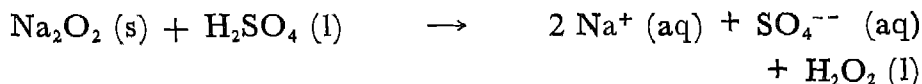


## 6.9 Hydrogen peroxide

Hydrogen peroxide,  $\text{H}_2\text{O}_2$ , is another important compound of hydrogen and oxygen. Its dilute solution is extensively used as a bleaching agent, particularly for delicate materials such as wool, silk, teeth and hair. In concentrated solutions, it is used as a fuel in rockets. In the laboratory it is used as an oxidizing agent.

### *Preparation of hydrogen peroxide*

Hydrogen peroxide is prepared by the action of a dilute acid upon peroxides of sodium or barium.



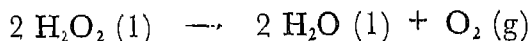
The reactants should be kept ice-cold because hydrogen peroxide decomposes at ordinary temperatures.

If barium peroxide and a calculated amount of dilute sulphuric acid are used, barium sulphate formed in the reaction is insoluble and can be removed by filtration, leaving a solution of hydrogen peroxide which can be concentrated.



### *Properties of hydrogen peroxide*

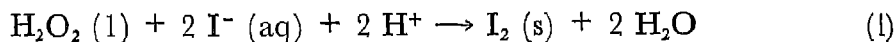
Hydrogen peroxide is a colourless liquid. As it is highly reactive, its dilute solutions are commonly used in the laboratory. Even in dilute solution it decomposes readily on warming to give water and oxygen.



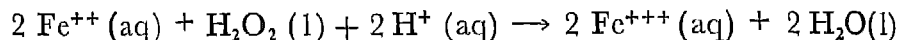
This decomposition is exothermic and is accelerated by alkalis, rough surfaces, finely divided metals, and by certain solids, particularly manganese dioxide

Hydrogen peroxide is an oxidizing agent. Some examples of its oxidizing action are given below :

- a) It liberates iodine from acidified potassium iodide solution.



- b) It oxidizes iron (II) salts to iron (III) salts

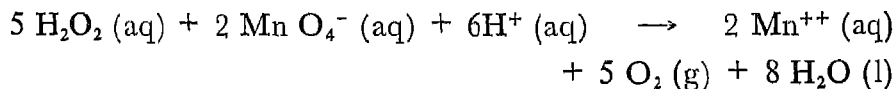


- c) It oxidizes black lead sulphide to white lead sulphate



Reaction (c) is used to restore oil paintings which have been darkened by the action of hydrogen sulphide in the atmosphere, the original lead pigments having changed to lead sulphide.

Hydrogen peroxide acts as a reducing agent with materials which are themselves strong oxidizing agents. For instance, it reacts with potassium permanganate in acid solution and is oxidized to oxygen.



### QUESTIONS

1. Write the equation for the preparation of water gas. Describe how this reaction is used for the commercial production of hydrogen.

2. Starting from hydrogen how do you obtain

- a) water
- b) hydrogen chloride
- c) hydrogen sulphide
- d) ammonia and
- e) sodium hydride ?

Write the equation for the reaction in each case.

3. Write three equations that illustrate the reducing properties of hydrogen. In each case point out the oxidizing agent.

4. Write the equation for the reaction between copper (II) oxide and hydrogen. Mention the (a) oxidizing agent (b) reducing agent (c) product of oxidation and (d) product of reduction involved in the reaction.

5. What products are formed when

- a) hydrogen is burnt in air or oxygen
- b) magnesium ribbon is heated in oxygen
- c) steam is passed over heated iron filings
- d) hydrogen is passed over heated copper (II) oxide ?

6. How is oxygen obtained from liquid air ?

On what principle is the isolation based ?

7. Write chemical equations mentioning conditions for the reaction of oxygen with

- a) magnesium
- b) sulphur
- c) nitrogen
- d) phosphorus
- e) carbon monoxide and
- f) iron (II) oxide

8. What is allotropy ? Name the allotrope of oxygen and mention some of its important uses.

9. Explain, with the help of chemical equations, the oxidizing properties of ozone.

10. Explain, in terms of electron loss and gain, the formation of (a) magnesium oxide and (b) sodium chloride.

11. What is natural water? What impurities are generally found in natural waters?
12. Why do we need to purify water for drinking purposes? Enumerate the different steps used in the purification of water
13. Which water would you call hard and which soft? Mention the disadvantages of using hard water.
14. How will you prove that water contains 1 part by volume of oxygen and 2 parts by volume of hydrogen?

Applying Avogadro's hypothesis deduce that the formula of water is  $\text{H}_2\text{O}$ .

15. What property of water makes it a good solvent for ionic substances?
16. Gaseous hydrogen chloride is a covalent compound but its aqueous solution contains hydrated hydrogen ions,  $\text{H}^+$  (aq). Explain.
17. Explain with examples the terms (a) deliquescence and (b) efflorescence.
18. What is the action of dilute sulphuric acid on a) sodium peroxide and b) barium peroxide?
19. Explain, with the help of chemical equations, the oxidizing action of hydrogen peroxide.
20. Tick off the correct answer in the following.

a) Oxidation is a process which involves

- i) gain of electrons
- ii) loss of electrons
- iii) both gain and loss of electrons

b) Reduction is a process which involves

- i) gain of electrons
- ii) loss of electrons
- iii) both gain and loss of electrons

## SOME CHEMICAL FAMILIES

Though there are more than a hundred elements discovered so far, the study of chemistry has been made easy because of similarities in the properties which enable us to group chemically related elements together. In this chapter we shall study three such groups of elements—the noble gases, the halogens and the alkali metals.

### 7.1 Noble gases

We know that the air around us contains besides nitrogen, oxygen and carbon dioxide, gases like argon, neon, krypton and xenon, argon being relatively more abundant. These, along with radon (which was isolated from radioactive materials) are called noble gases.

The noble gases have similar properties. They are, for instance, chemically non-reactive. It is for this reason they are called *inert* gases. Only recently a few compounds of some of these elements have been prepared, but they are relatively unstable and readily decompose to give back the rare gas.

The noble gases find extensive application in illumination. You might have seen 'neon' signs used for advertisements. These advertisement tubes consist of neon gas at low pressure mixed with other rare gases and mercury vapour. Electric bulbs are filled with argon gas. Helium is used in weather balloons and in dirigibles.

### 7.2 Halogens

We are familiar with some properties of chlorine, a constituent element of common salt. The elements bromine and iodine are also important in daily life. The compounds of bromine find use in medicine and photography. You may have heard of the bromide mixture administered to patients suffering from sleeplessness and of the bromide paper on which photographs are printed. Iodine is also familiar as the "tincture of iodine" used in dressing wounds. These three elements along with

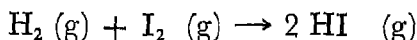
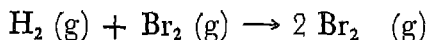
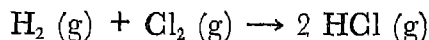
fluorine, and the radioactive element astatine, are called 'halogens' (producing sea-salt).

### *Properties of halogens*

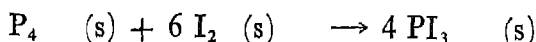
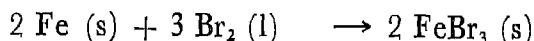
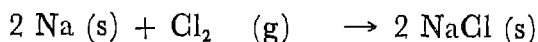
At ordinary temperatures, chlorine is a gas, bromine is a liquid and iodine, a solid. All the three elements are coloured. The colour intensifies from chlorine to iodine; chlorine is pale greenish-yellow, bromine is reddish-brown and iodine, dark purple or violet. The solubility of these elements in water decreases from chlorine to iodine. Chlorine is fairly soluble in water; bromine is moderately soluble and iodine very slightly so. They are more readily soluble in carbon tetrachloride. These elements have also characteristic odours—pungent and irritating.

### *Reactivity of halogens*

These elements react with hydrogen forming compounds known as hydrogen halides. Chlorine forms hydrogen chloride, bromine, hydrogen bromide and iodine, hydrogen iodide. Chlorine and hydrogen do not react in the dark, but do so, in diffuse light. The reaction occurs with explosive violence in bright sunlight. Bromine reacts with hydrogen less vigorously and only on heating. The reaction between iodine and hydrogen takes place rapidly only at high temperature.

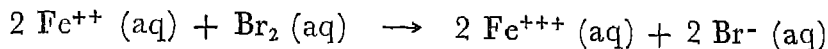
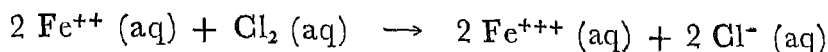


The three elements combine with metals like sodium, iron, aluminium and antimony and also with non-metals like phosphorus forming the corresponding halides.

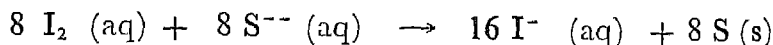
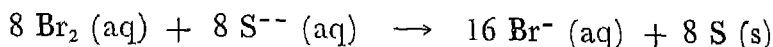
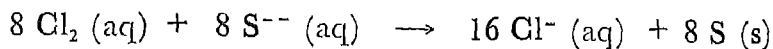


Chlorine is a strong oxidizing agent. It oxidizes iron (II) sulphate to iron (III) sulphate in aqueous solution at room temperature. Bromine oxidizes only in presence of an acid, whereas iodine cannot oxidize  $\text{Fe}^{++}$  to  $\text{Fe}^{+++}$ .

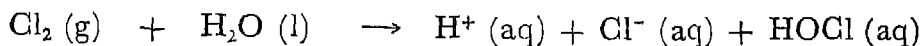




Chlorine oxidizes hydrogen sulphide to sulphur. Bromine and iodine also oxidize hydrogen sulphide, but less vigorously than chlorine.

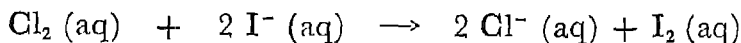
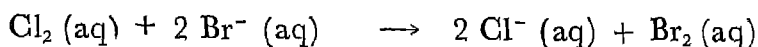


Chlorine acts as a bleaching agent in the presence of moisture. It bleaches many organic colouring matters like those in leaves, flowers, etc. Chlorine oxidizes these colouring matters readily to colourless products. In the absence of moisture it fails to act as bleaching agent. It appears the gas reacts with water to form hypochlorous acid, HOCl.

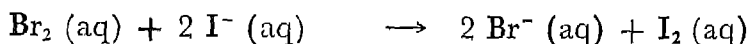


It is the hypochlorous acid which is the effective oxidizing agent in the decolourisation process. The bleaching action of bromine and iodine is extremely feeble.

When chlorine gas is passed through or when chlorine water is added to solutions of potassium bromide and potassium iodide, the chlorine oxidizes bromide to bromine and iodide to iodine.



Bromine, on the other hand, oxidizes only iodide to iodine.



Thus we find that of the three elements chlorine is the most reactive and iodine the least. Their reactivity is in the order

Chlorine > Bromine > Iodine

*Halides : chlorides, bromides and iodides*

The similarities among the elements extend to their compounds too. Hydrogen chloride, hydrogen bromide and hydrogen iodide are all gases

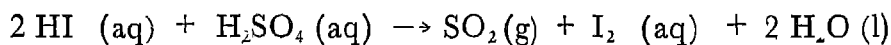
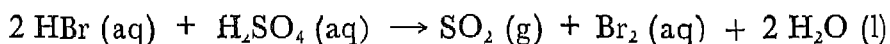
which, when dissolved in water, give acidic solutions. The halides of metals like potassium chloride, potassium bromide and potassium iodide are all colourless, crystalline, ionic solids and readily dissolve in water. They react with silver nitrate solution forming insoluble silver halides. Silver chloride is white, silver bromide pale yellow and silver iodide yellow. The silver halides show a gradation in their solubility in aqueous ammonia. For instance, silver chloride is soluble in aqueous ammonia, silver bromide less soluble and silver iodide least soluble.

When heated with concentrated sulphuric acid, the halides of metals like sodium and potassium give the corresponding hydrogen halides.



( $\text{X}^-$  refers to a  $\text{Cl}^-$ ,  $\text{Br}^-$  or  $\text{I}^-$  ion)

The hydrogen chloride formed is relatively stable. The hydrogen bromide and iodide are less stable and so are partially oxidized by sulphuric acid to bromine and iodine respectively.

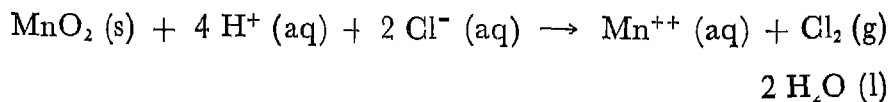


### *Sources of halogens*

Being reactive, these elements do not occur free in nature. Chlorine occurs as sodium chloride in sea water and in the mineral rock salt. Bromine occurs as potassium bromide in sea water. Sea water also contains iodides. These iodides are accumulated in sea weeds and thus sea weed is a source of iodine.

### *Preparation of halogens*

The halogens can be prepared by the oxidation of halides by means of oxidizing agents like manganese dioxide or potassium permanganate in the presence of aqueous acids.



Bromine and iodine are prepared by using bromides and iodides respectively. The halogens can also be prepared by electrolytic methods. However, in practice only chlorine and fluorine are prepared by these methods.

#### *Uses of halogens*

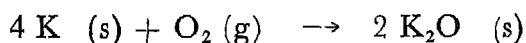
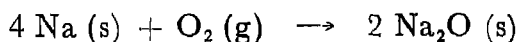
Chlorine is used to disinfect water for drinking purposes. It is also used to bleach wood-pulp (in paper industry), paper and textiles. Compounds of chlorine, bromine and iodine are used in a number of medicinal preparations. Compounds of bromine are used in the preparation of photographic chemicals.

### **7.3 The alkali metals**

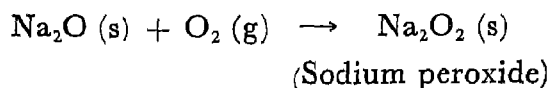
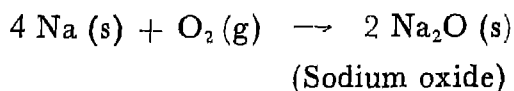
The elements lithium, sodium, potassium, rubidium, cesium and francium belong to a group of elements known as *alkali metals*. Among these we shall consider the properties of the two elements, sodium and potassium. These are silvery white metals with low melting points. They are ductile and are so soft that they can be cut with a knife. They are lighter than water and are good conductors of electricity.

#### *Reactivity of alkali metals*

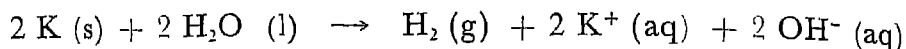
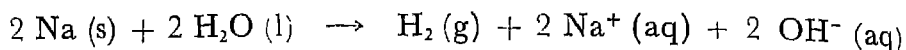
Both sodium and potassium are highly reactive. They tarnish readily in air due to the formation of a thin coating of the oxide.



Because of this reactivity, sodium and potassium are kept under kerosene. When heated in air they burn forming the oxides and peroxides.

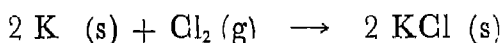
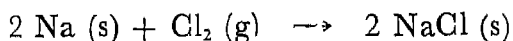


They also react vigorously with water to form hydrogen and the corresponding hydroxide. Potassium reacts more vigorously than sodium.



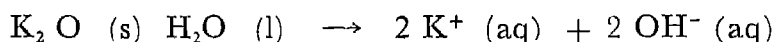
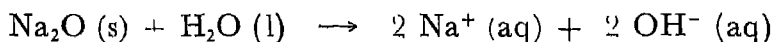
Large quantities of heat are released in these reactions.

When a piece of dry sodium or potassium is introduced into a jar of chlorine and heated, it reacts vigorously forming sodium or potassium chloride.



### *Alkali hydroxides*

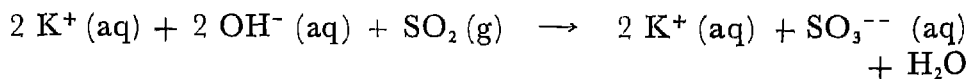
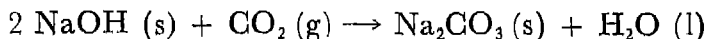
When dissolved in water, the oxides of sodium and potassium form the corresponding hydroxides.



The hydroxides of sodium and potassium are white solids and are also known as caustic soda and caustic potash respectively. They are deliquescent, that is, they absorb moisture from the atmosphere and form a concentrated solution. They dissolve in water producing much heat. The solutions are soapy to touch and turn red litmus blue. The hydroxides are highly corrosive. Their strong solutions cause blisters on the skin and damage clothing. They should, therefore, be handled with care. Bases like sodium hydroxide and potassium hydroxide, which are soluble in water, are called *alkalis*.

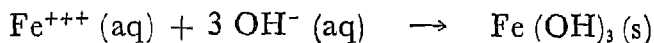
The hydroxides react with all acids to produce salts and water.

They also react with acidic oxides forming the corresponding salts and water. For instance, they react with carbon dioxide forming carbonates, and with sulphur dioxide forming sulphites.

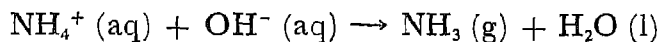


Alkali hydroxides react with many salt solutions forming another base and salt. If the newly formed base is insoluble in water, it is

precipitated. For example, iron (III) chloride solution reacts with sodium or potassium hydroxide to give a reddish-brown precipitate of iron (III) hydroxide.



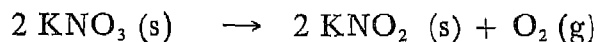
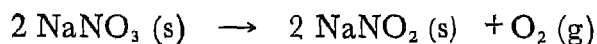
They also react with ammonium salts and liberate ammonia



#### *Salts of the alkali metals*

The salt of the alkali metals are white, ionic solids. Almost all of them are soluble in water.

When heated, the nitrates decompose giving nitrites and oxygen.



The other salts like chlorides and sulphates do not readily decompose on heating. The salts of these metals impart characteristic colour to the flame when heated in it. Sodium gives a golden yellow flame and potassium, lilac.

The other elements, lithium, rubidium, cesium and francium have also properties similar to those of sodium and potassium.

### **7.4 Electronic configuration and properties**

The study of the three groups of elements—noble gases, halogens and alkali metals reveals similarity in the properties of the elements in any group.

How can we account for this similarity? In chapter 4 we found that chemical properties of an element depend, to a large extent, on the valence electrons. Let us consider the electronic configuration of the noble gases, the halogens and the alkali metals (Table. 7.1).

TABLE 7.1

Electronic configuration of noble gases, alkali metals and halogens

NOBLE GASES			ALKALI METALS			HALOGENS		
Element	At. no.	Electronic configuration	Element	At. no.	Electronic configuration	Element	At. no.	Electronic configuration
Helium	2	2	Lithium	3	2,1	Fluorine	9	2,7
Neon	10	2,8	Sodium	11	2,8,1	Chlorine	17	2,8,7
Argon	18	2,8,8	Potassium	19	2,8,8,1	Bromine	35	2,8,18,7
Krypton	36	2,8,18,8	Rubidium	37	2,8,18,8,	Iodine	53	2,8,18,18,7
Xenon	54	2,8,18,18,8	Cesium	55	2,8,18,18,8,1	Astatine	85	2,8,18,32,18,7
Radon	86	2,8,18,32,18,8	Francium	87	2,8,18,32,18,8,1			

An examination of the table reveals that atoms of all the noble gases except helium (which contains two electrons) contain eight electrons in their outermost or valence shell. Likewise the atoms of the alkali metals have one valence electron each and those of the halogens, seven valence electrons. Thus, in any group the elements have similar electronic configuration. We may, therefore, conclude that the similarity in the configuration is responsible for the similarity in properties. In their chemical reactions the alkali metals tend to lose their valence electrons. Hence, they act as reducing agents. Whenever the halogens form ionic compounds they gain an electron and form the corresponding halide ion. Therefore, they act as oxidizing agents.

### QUESTIONS

1. If a substance A has the same number of electrons as another substance B, substance A and B are said to be isoelectronic. Pick out the isoelectronic groups from the following :  
K<sup>+</sup>, Na, Br<sup>-</sup>, Cl<sup>-</sup>, He, Ne, Ar.
2. Can burning sodium be put out by pouring water over it ? Explain.
3. Three jars labelled A, B and C contain different gases.

- (a) Burning sodium burns in jars A and C and deposits a white solid, but it is put out in jar B.
- (b) The gas in jar C is pungent smelling. The gases in jars A and B have no smell.
- (c) When potassium bromide solution is added to the three jars, it turns brown in jar C and is unchanged in jars A and B.
- Identify the gases
4. Answer whether the following statements are *TRUE* or *FALSE*.
- (a) The electronic configuration of the noble gases is taken as criteria for stability.
- (b) Chlorine atom has one electron more than the nearest noble gas.
- (c) The alkali metals and halogens tend to form ionic compounds.
- (d) The bond between chlorine atoms in a molecule of chlorine is ionic.
- (e) Iodine is the most reactive element among the halogens.
- (f) Hydrogen chloride is a polar molecule.
5. Fill up the blanks with suitable words or phrases.
- (a) When sodium reacts with water, much heat is.....
- (b) Chlorine combines with hydrogen to form gaseous.....
- (c) The white deposit formed when sodium burns in chlorine is.....
- (d) Among the halogens, bromine can displace....., but not ... ..from a solution of their salts.
- (e) When starch solution is added to iodine solution a..... ..colour is produced.
- (f) The alkali metal which gives a lilac flame is..... ..
- (g) .....is the halogen which can be purified by sublimation.
6. Match the appropriate item in column B against that in column A.

**A**

**B**

- |             |                                                           |
|-------------|-----------------------------------------------------------|
| 1. Bromine. | (a) Occurs as a component in common salt.                 |
| 2. Iodine   | (b) Is used as an antiseptic.                             |
| 3. Sodium   | (c) Is the only liquid nonmetal at ordinary temperatures. |
| 4. Argon.   | (d) Is used up in the respiration of living beings.       |
|             | (e) Is inert.                                             |

7. Which of the following represents the electronic configuration of the bromide ion,  $\text{Br}^-$  ?  
(a) 2, 8, 18, 1 (b) 2, 8, 18, 8 (c) 2, 8, 18, 7 (d) 2, 8, 18, 18.
8. Among the halogens, which would be a better *reducing* agent ?  
(a) chlorine (b) bromine (c) iodine.



## PERIODIC CLASSIFICATION OF ELEMENTS

In addition to the ninety and odd elements occurring in nature several other elements have been synthesised during the last three decades. Through a combination of two or more of these elements, over a million compounds are formed. The study of such a large number of substances will be almost impossible unless the elements are classified in some systematic way. Such a classification simplifies the study of chemistry and enables us to find similarities and variations in the properties of elements and their compounds. In this chapter we shall learn about the basis of such a classification.

### 8.1 The periodic law and the periodic table

In the previous chapter we learnt about three distinct families of elements—the noble gases, alkali metals and the halogens. We found that the similarity of properties makes it possible for the elements fluorine, chlorine, bromine and iodine to be grouped together. Likewise, the noble gases form a family of chemically similar elements and the alkali metals, another. Let us now see if there is any relation between the atomic numbers of similar elements and their properties.

Suppose the elements of these three families are arranged in the increasing order of their atomic numbers, as shown below.

<i>Noble gas</i>	<i>Alkali metal</i>	<i>Halogen</i>
He	Li	F "
(2)	(3)	(9)
Ne	Na	Cl
(10)	(11)	(17)
Ar	K	Br
(18)	(19)	(35)

The atomic number of an alkali metal is less by six than that of the following halogen element. Again, the atomic number of an alkali metal is greater by one than the preceding noble gas element. The atomic number of a halogen is one less than that of the noble gas in the next row. These facts indicate that noble gas elements have, as their immediate neighbours, the halogens on one side and the alkali metals on the other. By this arrangement the elements with similar properties fall in a group.

The existence of such regularities in the above arrangement suggests that the atomic number is a useful basis for the classification of elements.\* Such a classification was mainly due to the work of Henry Moseley, an English scientist. He showed that the atomic number of an element is a characteristic property of the element. When the elements are arranged in the increasing order of their atomic numbers, a well defined *periodicity* is observed in their properties. Periodicity means a repetition of events at regular intervals. The occurrence of the seasons is, for example, a periodic function of time. The variation of the properties of elements with their atomic numbers is stated in the form of the *periodic law* as follows :-

*The properties of elements are periodic functions of their atomic numbers.*

The arrangement of elements based on their properties is called the *periodic table* and is given in Fig. 8.1.

An examination of the periodic table shows that the elements fall into vertical '*columns*' or '*groups*' and horizontal '*rows*' or '*periods*'. The first row or period contains two elements-hydrogen and helium. The second and third rows contain eight elements each, while the succeeding rows contain increasing number of elements.

---

\* Before the elements were classified on the basis of their atomic numbers, several attempts were made to arrange them on the basis of their atomic weights. Among those who made such a classification were Dmitri Mendeleev, a Russian and Lothar Meyer, a German. Independently, in 1869, they classified the elements on the basis of their atomic weights. This was later discarded as it led to a number of discrepancies.

<sup>1</sup> <b>H</b> 1.008	PERIODIC TABLE OF ELEMENTS																<sup>2</sup> <b>He</b> 4.00
<sup>3</sup> <b>Li</b> 6.94	<sup>4</sup> <b>Be</b> 9.01															<sup>9</sup> <b>F</b> 19.0	<sup>10</sup> <b>Ne</b> 20.2
<sup>11</sup> <b>Na</b> 23.0	<sup>12</sup> <b>Mg</b> 24.3															<sup>17</sup> <b>Cl</b> 35.5	<sup>18</sup> <b>Ar</b> 39.9
<sup>19</sup> <b>K</b> 39.1	<sup>20</sup> <b>Ca</b> 40.1	<sup>21</sup> <b>Sc</b> 45.0	<sup>22</sup> <b>Ti</b> 47.9	<sup>23</sup> <b>V</b> 50.9	<sup>24</sup> <b>Cr</b> 52.0	<sup>25</sup> <b>Mn</b> 54.9	<sup>26</sup> <b>Fe</b> 55.8	<sup>27</sup> <b>Co</b> 58.9	<sup>28</sup> <b>Ni</b> 58.7	<sup>29</sup> <b>Cu</b> 63.5	<sup>30</sup> <b>Zn</b> 65.4	<sup>31</sup> <b>Ga</b> 69.7	<sup>32</sup> <b>Ge</b> 72.6	<sup>33</sup> <b>As</b> 74.9	<sup>34</sup> <b>Se</b> 79.0	<sup>35</sup> <b>Br</b> 79.9	<sup>36</sup> <b>Kr</b> 83.8
<sup>37</sup> <b>Rb</b> 85.5	<sup>38</sup> <b>Sr</b> 87.6	<sup>39</sup> <b>Y</b> 88.9	<sup>40</sup> <b>Zr</b> 91.2	<sup>41</sup> <b>Nb</b> 92.9	<sup>42</sup> <b>Mo</b> 95.9	<sup>43</sup> <b>Tc</b> (99)	<sup>44</sup> <b>Ru</b> 101.1	<sup>45</sup> <b>Rh</b> 102.9	<sup>46</sup> <b>Pd</b> 106.4	<sup>47</sup> <b>Ag</b> 107.9	<sup>48</sup> <b>Cd</b> 112.4	<sup>49</sup> <b>In</b> 114.8	<sup>50</sup> <b>Sn</b> 118.7	<sup>51</sup> <b>Sb</b> 121.8	<sup>52</sup> <b>Te</b> 127.6	<sup>53</sup> <b>I</b> 126.9	<sup>54</sup> <b>Xe</b> 131.3
<sup>55</sup> <b>Cs</b> 132.9	<sup>56</sup> <b>Ba</b> 137.3	<sup>57-71</sup> <b>See below</b>	<sup>72</sup> <b>Hf</b> 178.5	<sup>73</sup> <b>Ta</b> 180.9	<sup>74</sup> <b>W</b> 183.9	<sup>75</sup> <b>Re</b> 186.2	<sup>76</sup> <b>Os</b> 190.2	<sup>77</sup> <b>Ir</b> 192.2	<sup>78</sup> <b>Pt</b> 195.1	<sup>79</sup> <b>Au</b> 197.0	<sup>80</sup> <b>Hg</b> 200.6	<sup>81</sup> <b>Tl</b> 204.4	<sup>82</sup> <b>Pb</b> 207.2	<sup>83</sup> <b>Bi</b> 209.0	<sup>84</sup> <b>Po</b> (209)	<sup>85</sup> <b>At</b> (210)	<sup>86</sup> <b>Rn</b> (222)
<sup>87</sup> <b>Fr</b> (223)	<sup>88</sup> <b>Ra</b> (226)	<sup>89-</sup> <b>See below</b>															

<sup>57</sup> <b>La</b> 138.9	<sup>58</sup> <b>Ce</b> 140.1	<sup>59</sup> <b>Pr</b> 140.9	<sup>60</sup> <b>Nd</b> 144.2	<sup>61</sup> <b>Pm</b> (147)	<sup>62</sup> <b>Sm</b> 150.4	<sup>63</sup> <b>Eu</b> 152.0	<sup>64</sup> <b>Gd</b> 157.3	<sup>65</sup> <b>Tb</b> 158.9	<sup>66</sup> <b>Dy</b> 162.5	<sup>67</sup> <b>Ho</b> 164.9	<sup>68</sup> <b>Er</b> 167.3	<sup>69</sup> <b>Tm</b> 168.9	<sup>70</sup> <b>Yb</b> 173.0	<sup>71</sup> <b>Lu</b> 175.0
<sup>89</sup> <b>Ac</b> (227)	<sup>90</sup> <b>Th</b> 232.0	<sup>91</sup> <b>Pa</b> (231)	<sup>92</sup> <b>U</b> 238.0	<sup>93</sup> <b>Np</b> (237)	<sup>94</sup> <b>Pu</b> (242)	<sup>95</sup> <b>Am</b> (243)	<sup>96</sup> <b>Cm</b> (247)	<sup>97</sup> <b>Bk</b> (245)	<sup>98</sup> <b>Cf</b> (251)	<sup>99</sup> <b>Es</b> (254)	<sup>100</sup> <b>Fm</b> (253)	<sup>101</sup> <b>Md</b> (256)	<sup>102</sup> <b>No</b> (254)	<sup>103</sup> <b>Lw</b> (257)

The values given in parenthesis are the mass numbers of the isotopes with longest half lives.

Fig. 8.1 Periodic table of elements

## 8.2 Elements of the second row

Let us consider the second row which contains the elements lithium, beryllium, boron, carbon, nitrogen, oxygen, fluorine and neon. Table 8.1 summarises some properties such as melting and boiling points, metallic character and nature of oxides of these elements. We find that lithium has characteristic metallic properties; beryllium too is metallic, but less so than lithium. Boron is less metallic than lithium and beryllium, while carbon is mostly nonmetallic. Nitrogen, oxygen and fluorine are nonmetallic. Thus, there is a gradual change from metallic to nonmetallic character from lithium to fluorine, the last element, neon, being a gas.

This trend in properties is shown also by the oxides of these elements. There is a gradual change from a highly basic character to a highly acidic character.

## 8.3 Elements of the third row

The elements of this row are sodium, magnesium, aluminium, silicon, phosphorus, sulphur, chlorine and argon. These too show a trend in their properties, much like the elements of the second row. Table 8.2 summarises some properties of the elements of this row. Sodium is highly metallic; magnesium is only slightly less metallic than sodium. Aluminium is less so than magnesium. Silicon is even less metallic while phosphorus, sulphur and chlorine are nonmetals. Argon, like neon, is a noble gas. This trend in properties is shown also by the oxides of these elements.

The elements with similar properties fall in the same group or column of the periodic table. For example, the alkali metals are all in the same group. The halogens and the noble gases form separate groups. An examination of the electronic configuration of these elements shows that the number of valence electrons is the same for elements of similar properties.

In a given group the reactivity of the elements varies from element to element. For instance, among the alkali metals, the reactivity increases in the order lithium < sodium < potassium < rubidium < cesium. With the halogens, on the other hand, the reactivity decreases in the order fluorine > chlorine > bromine > iodine.

TABLE 8.1  
Some properties of the elements of the second row

Element	Lithium	Beryllium	Boron	Carbon	Nitrogen	Oxygen	Fluorine	Neon
Atomic number	3	4	5	6	7	8	9	10
Atomic weight	7.0	9.0	11.0	12.0	14.0	16.0	19.0	20.0
Melting point (°C)	180	1280	2300	3730	-210	-219	-223	-249
Boiling point (°C)	1326	2970	2500	4830	-196	-183	-188	-246
Metallic character*	Active metal	Less active metal	Few metallic properties	Intermediate between those of metals and non-metals	Nonmetallic properties	Nonmetallic properties	Strongly nonmetallic	—
Nature of oxide	Strongly alkaline	Weakly alkaline	Feebly acidic	Acidic	Strongly acidic	Decomposes water	—	—

\*Metallic character includes malleability, ductility, electrical conductivity, etc.

TABLE 8.2  
Some properties of the elements of the third row

Element	Sodium	Magnesium	Aluminium	Silicon	Phosphorus	Sulphur	Chlorine	Argon
Atomic number	11	12	13	14	15	16	17	18
Atomic weight	23.0	24.0	27.0	28.0	31.0	32.0	35.5	40.0
Appearance	Silvery white (soft)	Silvery white	Silvery white	Black solid	Yellow or red solid	Yellow solid	Greenish yellow gas	Colourless gas
Melting point (°C)	98	650	660	1410	44.2	119	-101	-189
Boiling point (°C)	889	1120	2327	2355	280	445	-34	-186
Nature of oxide	Highly basic	Basic	Feebly basic and acidic	Neutral	Acidic	Strongly acidic	Very strongly acidic	—

Thus, the periodic classification of elements helps us to understand clearly the large number of facts regarding the properties of elements and their compounds. The classification of elements into groups and rows is a direct result of the recurrence of properties at intervals when they are arranged in the increasing order of their atomic numbers. You will learn more about the periodic table and its usefulness in higher classes.

## QUESTIONS

1. State the periodic law.
2. Explain the trend in the properties of the elements of the second row of the periodic table.
3. How does the order of reactivity of the alkali metals differ from that of the halogens?
4. What factor is responsible for the similarity in the properties of the alkali metals?
5. Enumerate the trends in the nature of the oxides of the third row elements.
6. Lithium forms the following compounds.

Lithium oxide,  $\text{Li}_2\text{O}$ ; lithium hydroxide,  $\text{LiOH}$ ; lithium sulphide,  $\text{Li}_2\text{S}$ .

Write the formulae of the corresponding sodium and potassium compounds.

7. Arrange the acids formed from  $\text{SO}_3$ ,  $\text{Cl}_2\text{O}_7$  and  $\text{P}_2\text{O}_5$  in the increasing order of acidity.

